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Foreword

Based largely on the pattern of materials used in the schools of USSR where science is taught as separate disciplines from the middle stage, the first edition of Chemistry Part II was published in 1969. After getting the feedback from the users and others the content of the book has been completely reorganised and is now appearing in this revised edition.

This book like the others in the series emphasises the conceptual understanding and reasoning rather than retention and reproduction of facts. The curriculum on the basis of which this book has been written has a practical bias and the student is made to gain knowledge through experiements conducted by teachers and pupils.

This book will now be used in the 3rd year of the selected middle schools under the pilot project for the improvement of science education in schools. With the feedback obtained from the various States, subsequent edition of the book would be further improved.

The first textual materials were developed by a team of writers consisting of Dr M. C. Pant, Shri N. K. Sanyal, Dr. B. D. Atreya and Dr. C. Radhakrishnan. Dr. S. V. Balezin, Dr. Y.I. Naumov and Dr. L. V Levchuk acted as our UNESCO Consultants. In further revising the book Dr. K. C. Mathur and Shri K. S. Bhandari have also assisted. In preparing the present reprint edition considerable effort has been put in by Shri N.K. Sanyal, Dr. M C. Pant Dr. K. M. Pant, Dr. P.S. Verma (Pool Officer, C.S.I.R) Shri B.L. Pandit and Dr. L.V. Levchuk, UNESCO Consultant. Our appreciation and grateful thanks are due to all of them and to the Chemistry Group in our Department of Science Education.

Suggestions for further improvement of the book are welcome and would be gratefully acknowledged.

New Delhi
December 1971

S. V. C. Aiya
Director

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CHAPTER I

Important Classes of Inorganic Compounds

YOU have learnt that a wide variety of substances are found which may be simple substances or compounds or a mixture of them. Some substances are found in nature as such. Others are prepared from such sources. There are some substances which are produced from living organisms like oils, fats, sugar, starch, wood and coal etc. These are called *organic* substances, and you will learn more about them later. The others which are not organic, are called ***inorganic*** substances. These substances may be of several types like oxides, acids, bases and salts.

We shall study something more about these important classes of inorganic compounds.

OXIDES

You know that oxides are binary compounds of oxygen and another element. Many oxides can be prepared by the direct combination of oxygen with a metal or a non-metal. Some oxides like carbon dioxide, sulphur dioxide, water etc. occur in nature

We shall first study the oxides of metals.

1. Basic Oxides (Metallic Oxides)

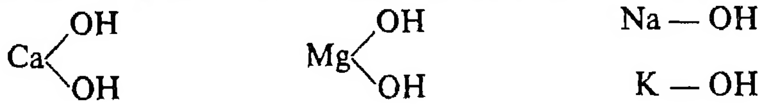
Experiment : Take a small quantity of quicklime (calcium oxide) in a dry test tube and observe the state and colour of the substance. Then carefully pour a few millilitres of water in the test tube Observe what happens.

Decant a few clear drops of this solution on a watch glass and add a drop of phenolphthalein solution. Explain what happens

This experiment may be repeated with magnesium oxide or sodium oxide. The chemical reactions are represented as follows :

CaO	+	H_2O	=	$Ca(OH)_2$
Calcium Oxide		Water		Calcium hydroxide
MgO	+	H_2O	=	$Mg(OH)_2$
Magnesium oxide		Water		Magnesium hydroxide
Na_2O	+	H_2O	=	$2NaOH$
Sodium oxide		Water		Sodium hydroxide

All these substances form hydroxides. They have a common structure of having one or more hydroxyl (OH) groups. For example :



Experiment : Repeat the previous experiment using copper oxide and iron oxide separately. Observe what happens in these cases

Unlike sodium or magnesium oxides, copper and iron oxides do not react directly with water to form hydroxides.

Experiment : Take a porcelain dish and place in it a pinch of magnesium oxide. Then add carefully 1-2 millilitres of dilute sulphuric acid. Observe what happens. Take a clear drop of the solution by a dropper and evaporate carefully on a watch-glass. Is any residue left ?

Repeat the same experiment using about 0.5 g of copper oxide. Does the oxide dissolve in dilute sulphuric acid readily ? Is there any change of colour ? Warm the dish above the flame of a spirit lamp and now observe what happens.

Take a few drops of a clear blue solution and evaporate on a watch glass. Observe the blue crystals obtained.

In the above experiments both magnesium oxide and copper oxide dissolve in dilute sulphuric acid to give salts of sulphuric acid which are called *sulphates*. The reactions can be expressed as follows :

MgO	+	H_2SO_4	=	$MgSO_4$	+	H_2O
Magnesium oxide		Sulphuric acid		Magnesium sulphate (white)		Water

CuO	+	H_2SO_4	=	$CuSO_4$	+	H_2O
Copper oxide		Sulphuric acid		Copper sulphate (blue)		Water

Another experiment for studying the interaction of oxides of metals and acids may be as follows:

Experiment : Take a porcelain dish and place in it a pinch of ferric oxide. Then add carefully a few millilitres of dilute hydrochloric acid. Observe what happens. Take a clear drop of the solution by a dropper and evaporate carefully on a watch glass. Is any residue left ? Now warm the content of the porcelain dish and observe what happens.

Take a few drops of the clear solution and evaporate on a watch glass. Observe the colour of the crystals obtained.

The reaction is expressed by the following equation :

$6\,HCl$	+	Fe_2O_3	=	$2FeCl_3$	+	$3H_2O$
Hydrochloric acid		Ferric oxide		Ferric chloride		Water

Similarly by the interaction of magnesium oxide (MgO) with nitric acid, a colourless solution of magnesium nitrate is obtained. The equation of the reaction is as follows :

$2HNO_3$	+	MgO	=	$Mg(NO_3)_2$	+	H_2O
Nitric acid		Magnesium oxide		Magnesium nitrate		Water

You have thus seen that oxides of metals react with acids to form salt and water and some oxides combine with water to form hydroxides.

Oxides which form salt and water with acids are called basic oxides.

Sodium oxide, potassium oxide, calcium oxide (quicklime), barium oxide, magnesium oxide, copper oxide, iron oxide are examples of basic oxides. Basic oxides are obtained *only* from metals. They *do not react* with bases.

Some of the basic oxides like sodium oxide, calcium oxide, magnesium oxide readily dissolve in water to form hydroxides.

If a metal forms two oxides, the name of that oxide in which the metal has a lower valency has the suffix **'ous'** and the name of that oxide in which the metal has a higher valency has the suffix **'ic'**. For example

Cu_2O	—	*Cuprous oxide*	Valency 1
CuO	—	*Cupric oxide*	Valency 2
FeO	—	*Ferrous oxide*	Valency 2
Fe_2O_3	—	*Ferric oxide*	Valency 3

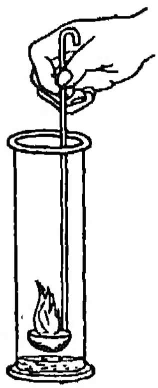


FIG. 1. Burning sulphur in oxygen

2. Acidic Oxides (Non-Metallic Oxides)

Experiment : Take a small amount of sulphur in a deflagrating spoon provided with a cardboard lid. Ignite the sulphur and plunge it carefully in a gas jar containing a few drops of water (Fig. 1). Close the cardboard cover. After the sulphur is extinguished, close the jar with a lid and shake it. Then pour two drops of litmus solution and observe what happens.

This experiment may be repeated by using carbon (charcoal) and phosphorus.

Demonstration : Fit an apparatus to prepare a continuous supply of carbon dioxide. Take a test tube containing a saturated solution of quicklime (lime water) Pour a drop of phenolphthalein. What colour does it indicate ?

Now pass a slow and steady stream of carbon dioxide bubbles. Watch if there is any change of colour. How would you explain it ?

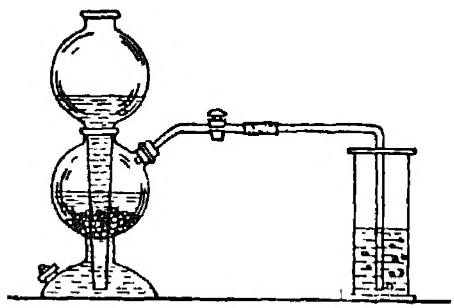


FIG. 2. Passing carbon dioxide in lime water

The above reactions can be explained by the following equations:

(*i*) $S + O_2 = SO_2$

$SO_2 + H_2O = H_2SO_3$

(*ii*) $CO_2 + H_2O = H_2CO_3$
$H_2CO_3 + Ca(OH)_2 = CaCO_3 + 2H_2O$

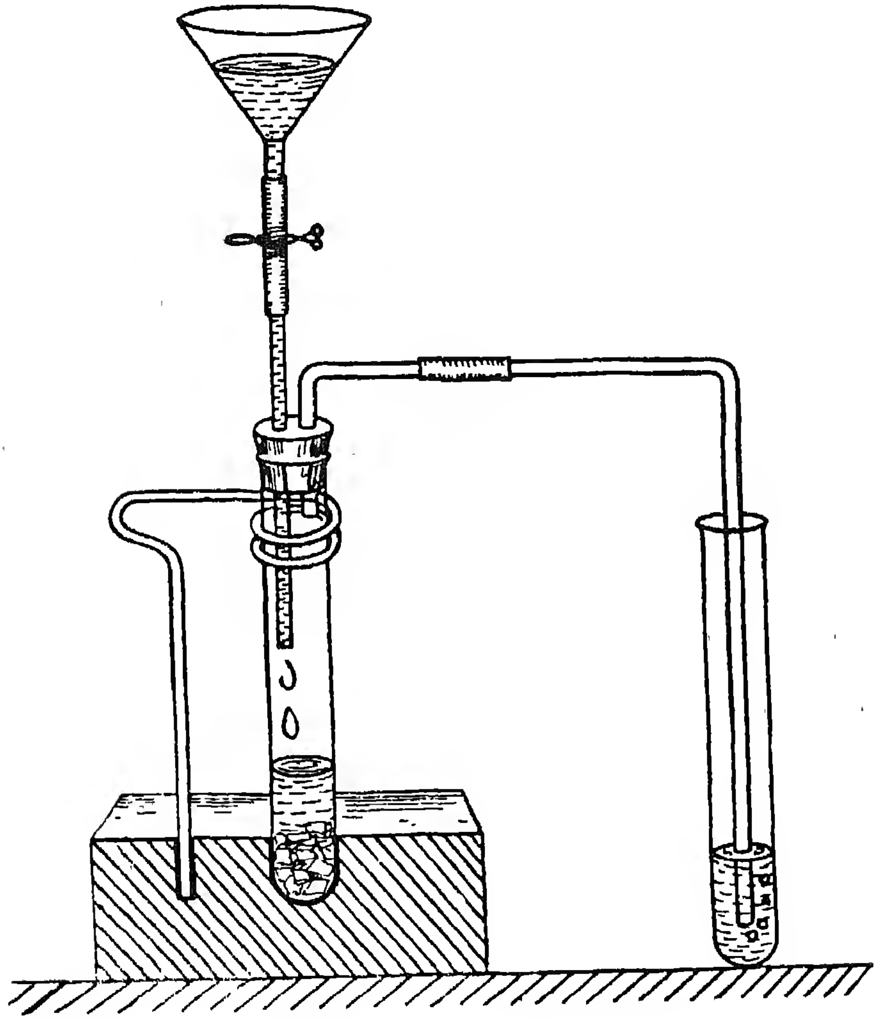


FIG. 3. Passing carbon dioxide in lime water (Using the kit)

This experiment could be used with any other acidic oxide like SO_2 or P_2O_5 and other base $Ba(OH)_2$.

These reactions can be expressed by the following equations :

$$SO_2 + H_2O = H_2SO_3$$
$$H_2SO_3 + Ba(OH)_2 = BaSO_3 + 2H_2O.$$

Thus non-metallic oxides dissolve in water to form acids which can react with bases to form salt and water.

Oxides which form salt and water with bases are called acidic oxides.

Acidic oxides are also called *anhydrides of acids* because by removal of water from acids these oxides can be obtained.

H_2SO_3 Sulphurous acid	=	H_2O Water	+	SO_2 Sulphurous anhydride
H_2CO_3 Carbonic acid	=	H_2O Water	+	CO_2 Carbonic anhydride
H_2SiO_3 Silicic acid	=	H_2O Water	+	SiO_2 Silicic anhydride

Although acidic oxides are mostly obtained from non-metals, some metals also form acidic oxides. In such cases these metals exhibit a higher valency in the acidic oxide compared to the basic oxide. For example :

Cr_2O_3	—	Chromous oxide (basic—valency 3)
CrO_3	—	Chromic anhydride (acidic—valency 6)
MnO	—	Manganous oxide (basic—valency 2)
Mn_2O_7	—	Permanganic anhydride (acidic—valency 7)

The names of anhydrides are derived from the names of the acids from which they are obtained. For example :

CO_2	—	Carbonic anhydride	H_2CO_3	—	Carbonic acid
SO_2	—	Sulphurous ,,	H_2SO_3	—	Sulphurous ,,
SO_3	—	Sulphuric ,,	H_2SO_4	—	Sulphuric ,,
P_2O_5	—	Phosphoric ,,	H_3PO_4	—	Phosphoric ,,
N_2O_5	—	Nitric ,,	HNO_3	—	Nitric ,,
SiO_2	—	Silicic ,,	H_2SiO_3	—	Silicic ,,

Questions

1. What are basic oxides ? Write the formulae of four basic oxides giving the valency of the elements present in them.
2. What are acidic oxides ? Write the formulae of four acidic oxides.
3. Write the formulae of the anhydrides of silicic acid (H_2SiO_3) and phosphoric acid (H_3PO_4).
4. What will be the formula of the hydroxides of the following oxides : BaO, K_2O.

3. Chemical Properties of Basic Oxides

(*a*) *Interaction of basic oxides with water* (*reaction of hydration*)

You know that only some basic oxides take part in reaction of hydration resulting in the formation of soluble bases (alkalis).

You have already learnt the reaction of hydration of quicklime which is industrially known as slaking (quenching of lime). Barium oxide also undergoes such reaction of hydration :

$$\underset{\text{Barium oxide}}{BaO} + \underset{\text{Water}}{H_2O} = \underset{\text{Barium hydroxide}}{Ba(OH)_2}$$

Similarly sodium oxide and potassium oxide react with water and produce soluble bases :

$$Na_2O + H_2O = 2NaOH$$
$$K_2O + H_2O = 2KOH$$

Many basic oxides are insoluble and do not react with water, e.g., copper oxide (CuO), iron oxides (FeO and Fe_2O_3).

(*b*) *Interaction of basic oxides with acids*

When magnesium oxide interacts with hydrochloric acid, the salt magnesium chloride and water are obtained. Similarly ferrous oxide gives the salt, ferrous sulphate and water by interaction with sulphuric acid. Basic oxides react with acids to form salt and water :

$$MgO + 2HCl = MgCl_2 \quad H_2O$$
$$FeO + H_2SO_4 = FeSO_4 + H_2O$$
$$\text{Basic oxide} + \text{Acid} = \text{Salt} + \text{Water}$$

Interaction with acids is one of the important properties of basic oxides.

Both soluble and insoluble basic oxides take part in the reaction with acids.

(*c*) *Reaction of combination of basic and acidic oxides*

Demonstration : Make a loop at the tip of a deflagrating spoon. Clean it by dipping it into a dish containing dilute hydrochloric acid and heating it over the flame of a spirit lamp. Repeat it two or three times. Wash with water and dry it over the flame.Touch the loop with solid borax and heat it on the flame till a clear bead of boron oxide (B_2O_3) is formed. Touch the hot bead with a little cupric oxide and heat again strongly (Fig. 4). What do you observe ? What is the nature of the new substance formed.

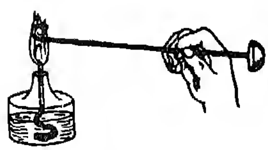


FIG 4 Heating borax bead on a flame

The basic copper oxide has formed a coloured salt with acidic boron oxide as follows :

$$3CuO \quad + \quad B_2O_3 \qquad\qquad Cu_3(BO_3)_2$$

Quicklime (CaO) and silica (SiO_2) or sand react together on heating strongly to give a glassy mass of calcium silicate, a salt :

$$\underset{\text{Calcium oxide}}{CaO} \quad + \quad \underset{\text{Silicon dioxide}}{SiO_2} \quad \overset{\text{Heat}}{=} \quad \underset{\text{Calcium silicate}}{CaSiO_3}$$

This reaction is used for the preparation of glass.

Magnesium oxide similarly reacts energetically with solid sulphuric anhydride to give magnesium sulphate. A lot of heat is liberated in this reaction :

$$\underset{\text{Magnesium oxide}}{MgO} \quad + \quad \underset{\text{Sulphuric anhydride}}{SO_3} \qquad\qquad \underset{\text{Magnesium sulphate}}{MgSO_4} \quad + \quad \text{heat}$$

Interaction of acidic oxides (anhydrides of acids) and basic oxides to form salts is another characteristic chemical property of these oxides.

Questions

1. Write the equation of the hydration of calcium oxide and name the substance obtained.
2. How many kilograms of water are necessary to complete interaction with 280 kg of CaO ? **(Ans. 90 kg)**
3. State whether the following statements are correct or not :
   (*a*) Ferric oxide contains a higher percentage of oxygen than ferrous oxide.
   (*b*) Ferric oxide contains a lower percentage of iron than ferrous oxide.
   (*c*) Cupric oxide contains higher percentage of oxygen than cuprous oxide.
   (*d*) Cupric oxide contains a lower percentage of copper than cuprous oxide.
4. What happens when the following bases are dehydrated ? Give equations for the reactions involved.
   $Zn(OH)_2$, $Al(OH)_3$, $Cr(OH)_3$
5. What quantity of salt would be obtained by the action of nitric acid on 40 gm of ferric oxide ? **(Ans. 90 g)**
6. If 52.2 g of barium nitrate are obtained by the interaction of nitric acid and barium oxide (BaO), how many grams of the latter were taken ? **(Ans. 30.6 g)**
7. How are the following obtained ?
   BaO, ZnO, CuO.
   Write equations for the reactions involved.
8. Give equations of the reactions by which the following conversions could be effected :
   (*a*) $CaO \rightarrow Ca(NO_3)_2$
   (*b*) $CuO \rightarrow CuSO_4 \rightarrow Cu(OH)_2 \rightarrow CuO$

9. Write the equation for the reaction of :

   (*a*) Sulphuric anhydride with ferrous oxide.

   (*b*) Phosphoric anhydride with magnesium oxide.

4. Chemical Properties of Acidic Oxides

(*a*) *Interaction of acidic oxides with water (Reaction of hydration)*

You have already learnt that **carbon dioxide, sulphurous anhydride** (sulphur dioxide) **and** sulphuric anhydride **(sulphur trioxide) dissolve in water to** give corresponding acids. This is a ***reaction of hydration.*** These reactions can be expressed as follows :

CO_2 Carbonic anhydride	$+$	H_2O Water	$=$	H_2CO_3 Carbonic acid (Hydrate of carbon dioxide)
SO_2 Sulphurous anhydride	$+$	H_2O Water	$=$	H_2SO_3 Sulphurous acid (Hydrate of sulphur dioxide)
SO_3 Sulphuric anhydride	$+$	H_2O Water	$=$	H_2SO_4 Sulphuric acid (Hydrate of sulphur trioxide)

While many acidic oxides undergo the reaction of hydration, silica (silicic anhydride), SiO_2, does not. Hence it is not possible to get H_2SiO_3 (silicic acid) by the reaction of hydration.

(*b*) *Interaction of acidic oxides with bases*

Demonstration : Fill a big round bottom flask with carbon dioxide (Fig. 5). Ensure that it is full of carbon dioxide. Now cut 2 to 3 g of caustic soda into small pieces of the size of a rice grain. Introduce these pieces into the flask. Close the flask with a rubber stopper, fitted with a delivery tube, with a jet on the inner side and a rubber

tube with a pinch cock at the other end (Fig. 6). Keep the pinch cock closed before inserting the stopper into the mouth of the flask.

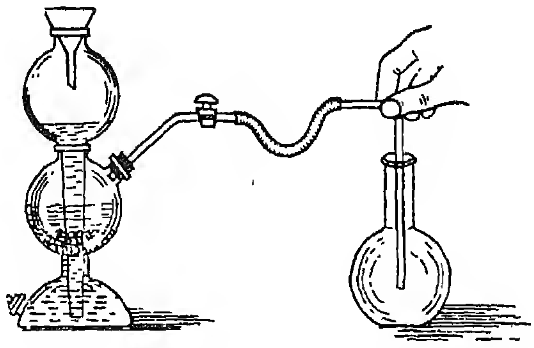


FIG. 5. Filling a flask with carbon dioxide

Observe what happens. See carefully the pieces of caustic soda and the walls of the flask. Touch the bottom of the flask by the palm of your hand. Invert the flask in a trough of water and open the pinch cock under water (Fig. 7). Observe what happens and explain

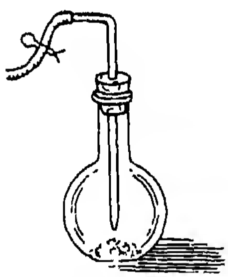


FIG. 6. Flask of carbon dioxide with caustic soda and a delivery tube closed with a pinch cock

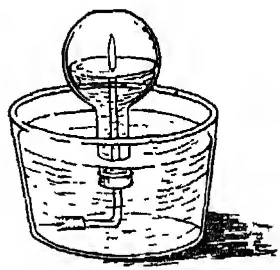


FIG. 7. Opening the delivery tube under water

The gaseous acidic oxide, carbonic anhydride, reacted with the basic caustic soda. This can be represented by the following equation.

$2NaOH$	+	CO_2	=	Na_2CO_3	+	H_2O
Base (Sodium hydroxide)		Anhydride (Carbonic)		Salt (Sodium carbonate)		Water

Similarly salts are also formed with different anhydrides of acid (acidic oxides) when they react with soluble bases (alkalis). The following are some examples :

1. $SO_2 + 2KOH = K_2SO_3 + H_2O$

SO_2	+	$2KOH$	=	K_2SO_3	+	H_2O
Sulphur dioxide (sulphurous anhydride)		Potassium hydroxide (Base)		Potassium sulphite (Salt)		Water

2. $SO_3 + 2NaOH = Na_2SO_4 + H_2O$

SO_3	+	$2NaOH$	=	Na_2SO_4	+	H_2O
Sulphur trioxide (sulphuric anhydride)		Sodium hydroxide (Base)		Sodium sulphate (Salt)		Water

3. $SiO_2 + Ba(OH)_2 = BaSiO_3 + H_2O$

SiO_2	+	$Ba(OH)_2$	=	$BaSiO_3$	+	H_2O
Silicon dioxide (Silicic anhydride)		Barium hydroxide (Base)		Barium silicate (Salt)		Water

4. $P_2O_5 + 3Ca(OH)_2 = Ca_3(PO_4)_2 + 3H_2O$

P_2O_5	+	$3Ca(OH)_2$	=	$Ca_3(PO_4)_2$	+	$3H_2O$
Phosphoric pentoxide (phosphoric anhydride)		Calcium hydroxide (Base)		Calcium phosphate (Salt)		Water

The ability of acidic oxides to interact with bases is one of their characteristic properties. This reaction *produces a salt and water*. Many of these reactions are used for the preparation of salts.

(*c*) *Interaction of acidic oxide with basic oxide*

You have already seen the reaction of acidic oxides with basic oxides. For example : sulphur trioxide or nitrogen pentoxide would directly react with sodium or potassium oxide to give salts as follows :

1. $SO_3 + Na_2O = Na_2SO_4$

SO_3	+	Na_2O	=	Na_2SO_4
Sulphur trioxide (Acidic (oxide)		Sodium oxide (Basic oxide)		Sodium sulphate (Salt)

2.	N_2O_5	+	K_2O	=	$2KNO_3$
	Nitric anhydride (Acidic oxide)		Potassium oxide (Basic oxide)		Potassium nitrate (Salt)

Questions

1. Write the equations of the reactions by which the following conversions may be effected :

   (*a*) $C \rightarrow CO_2 \rightarrow BaCO_3$

   (*b*) $SO_3 \rightarrow H_2SO_4 \rightarrow K_2SO_3$

2. Write the formulae of the anhydrides of the following acids :

   (*a*) Nitric ; (*b*) Sulphuric ;

   (*c*) Phosphoric ; (*d*) Sulphurous.

3. What is the characteristic chemical property of an acidic oxide ? Give the equations of the reaction that represent it.

4. What are the substances obtained by the hydration of acidic oxides ? Give equations to illustrate your answer.

5. Given phosphoric acid, what are the substances you need for obtaining the following salts :

   (*a*) $Ba_3(PO_4)_2$ (*b*) $AlPO_4$ (*c*) $Cu_3(PO_4)_2$

6. How many kilograms of salt will be obtained by the action of 120 kg of caustic soda with carbonic adhydride.

   (Ans. 159 kg)

7. How many quintals of sulphuric acid could be obtained from 40 quintals of sulphuric anhydride ? **(Ans. 49 quintals)**

8. Fill in the blank spaces with the type of oxide (acidic or basic) :

   (*a*) BaO______________

   (*b*) P_2O_5______________

   (*c*) SO_3______________

(*d*) CO_2____________

(*e*) CO____________

(*f*) H_2O____________

How will you verify whether your assessment is correct or not ?

Practical Work No. 1

Interaction of basic oxides with acids. Preparation of blue vitriol by the interaction of copper oxide with dilute sulphuric acid.

Apparatus required : Spirit lamp, metallic stand with ring, porcelain triangle for funnel, asbestos gauze, spoon, porcelain evaporating dish, beaker, funnel, and filter paper.

Reagents : Copper oxide and dilute sulphuric acid.

Procedure : Pour half a test-tube full of sulphuric acid into the porcelain dish. Heat it on the asbestos gauze and add copper oxide in small portions. Add fresh portion only after the previous one is dissolved. See if any undissolved copper oxide is left.

Put the folded filter paper into the funnel and moisten it with water. Filter the hot solution from the excess copper oxide into the beaker. Pour out a portion of the solution into the porcelain dish. Cover the dish with the funnel (why ?). Evaporate the solution till crystals of copper sulphate are obtained.

Pour out the solution of the salt into a bottle labelled "$CuSO_4$ Solution". Dry the crystals between folds of blotting or filter paper and transfer them into a wide mouthed bottle labelled "$CuSO_4 . 5H_2O$".

Clean your apparatus and working place. Make a report of your work.

Acids

You already know that acids are compounds which change the blue litmus to red and that their molecules contain atoms of hydrogen combined with an acid remnant (radical). They take part in the reaction of neutralization with bases forming salt and water.

Let us study something more about this important class of compounds.

5. Chemical Properties of Acids

(a) *Action of acids on indicators*

Experiment : In four test tubes, take 2-3 ml each of dilute hydrochloric acid, sulphuric acid, nitric acid and phosphoric acid. Add into each of these tubes a few drops of a violet solution of litmus. Observe the changes taking place.

Repeat the experiment by using a solution of methyl orange in place of litmus and make your observations.

Repeat using phenolphthalein as the indicator. What do you observe ?

Make a table in your note book as shown below and record your observations.

Acids	*Change in colour of indicator*		
	Litmus	*Methyl orange*	*Phenolphthalein*
Sulphuric			
Hydrochloric			
Nitric			
Phosphoric			

Acids change the colour of the indicators differently. The violet colour of litmus is changed to red. The orange colour of methyl orange changes into red by the action of acids. But this red colour has a different tinge from the red colour obtained in the case of litmus. With phenolphthalein, the colourless solution does not acquire any colour by the action of acids.

NOTE Study the precautions to be taken in working with acids and alkalies as given under 'Uses of acids' at the end of this section.

(*b*) *Interaction of acids with metals*

Experiment : Take about 2 ml of dilute sulphuric acid in a test tube and put 2-3 granules of zinc. Observe what happens. Close the test tube with a stopper carrying a delivery tube (see Fig. 8). Bring a lighted match stick near the opening of the delivery tube and see what happens. After the reaction is over, decant the liquid and evaporate a few drops of the filtrate on watch glass or a beaker cover as you did earlier. What do you observe ? What does it look like ? Is there any residue ?

FIG. 8. Evolution of hydrogen in a test tube

Repeat the experiment by using hydrochloric acid in place of sulphuric acid and substitute zinc by magnesium. Test the gas evolved. Write the equation of the reaction.

These experiments show that as a result of interaction of zinc and magnesium with acids, hydrogen is liberated and salts are obtained. The equations for the reactions are :

H_2SO_4	+	Zn	=	$ZnSO_4$	+	$H_2\uparrow$
Sulphuric acid (in solution)		Zinc (solid)		Zinc sulphate (in solution)		Hydrogen (gas)
2HCl	+	Mg	=	$MgCl_2$	+	$H_2\uparrow$
Hydrochloric acid (in solution)		Magnesium (solid)		Magnesium chloride (in solution)		Hydrogen (gas)

In the same way aluminium, iron and some other metals react with acids to give salts and hydrogen.

What kind of reactions are these ?

All these reactions belong to the reactions of displacement, since the hydrogen atoms are displaced by the metal atoms from the acid molecules.

Do all the metals react with solutions of hydrochloric and sulphuric acid in the same way as in the above case to give salts and hydrogen ? Let us find this out by an experiment.

Experiment : Take two test tubes containing about 2 ml each of dilute sulphuric acid and dilute hydrochloric acid respectively. Put a few copper turnings into each. Observe what happens.

Do you find any effervescence taking place ? Try to burn the gases evolved, if any, at the mouth of the delivery tube. What do you observe ? Does it burn in each case ?

You must have observed that under these conditions copper does not react with dilute H_2SO_4 and dilute HCl, with the evolution of hydrogen, as it happens in the case of zinc, magnesium and some other metals. Chemists have found out from experiments that under similar conditions mercury, silver, gold and platinum also behave like copper with these acids.

Let us see how copper reacts with nitric acid.

Experiment : Take a few millilitres of concentrated nitric acid in a test tube and put some copper turnings. Observe the colour of the solution in the test tube and note the colour of the gas evolved, if any. Test the gas with a burning splinter. Does it burn ?

Evaporate a few drops of the solution that is left behind on a watch glass. Observe the nature and colour of the residue.

In this case, instead of hydrogen, a reddish brown gas, nitrogen dioxide is evolved. The green coloured solution contains a salt, copper nitrate, which is obtained as a green crystalline solid on evaporation.

The reaction is expressed as follows :

$$4\,HNO_2 + Cu \quad = \quad Cu(NO_3)_2 \quad + \quad 2NO_2\uparrow \quad + \quad 2H_2O$$

Similar experiments with other metals have shown that nitric acid reacts with almost all the metals giving hydrogen or gaseous oxides of nitrogen.

(*c*) *Interaction of acids and oxides of metals*

Experiment : Take 3 test tubes each containing about 3 ml of hydrochloric acid and another 3 test tubes containing sulphuric acid.

Add a pinch of magnesium oxide, copper oxide and ferric oxide separately in each acid by turns and watch what happens. If necessary, warm the test tubes a little.

Evaporate separately about 5 ml of clear liquid of each to dryness. What do you get in each case ?

Make a table as follows in your note book and record your observations :

Acid	*Magnesium oxide* MgO	*Copper oxide* CuO	*Ferric oxide* Fe_2O_3
HCl			
H_2SO_4			

Write the equations of these reactions.

You have studied in section 3 that acids react with oxides of metals to form salt and water; this is thus a characteristic chemical property of both these classes of compounds.

It is possible to obtain a desired salt by choosing the proper acid and metallic oxide which react with each other. For example, by dissolving zinc oxide in nitric acid one can obtain zinc nitrate :

$$2HNO_3 + ZnO = Zn(NO_3)_2 + H_2O$$

(*d*) *Interaction of acids with bases* (*Reaction of neutralization*)

You know that salt and water are obtained from the reactions of neutralization.

Experiment : Take a beaker containing a solution of sulphuric acid. Add to it a few drops of phenolphthalein. Now add by a dropper a solution of alkali (caustic soda) a little at a time in the beginning, and drop by drop towards the end, till the colour of the solution turns pink by the addition of a single drop. Now take a few drops of the solution and evaporate on a watch glass. Observe the residue left on the watch glass.

Write the equations of this reaction.

Repeat the previous experiment using barium hydroxide in place of sodium hydroxide. What do you observe ? Is any residue left ?

The salt which is obtained in this reaction is insoluble and hence it precipitates :

H_2SO_4	+	$Ba(OH)_2$	=	$BaSO_4\downarrow$	+	$2H_2O$
Sulphuric acid		Barium hydroxide		Barium sulphate (insoluble precipitate)		Water

So far we have seen the action of acids on soluble bases. Let us see the action of acids on copper hydroxide, an insoluble base.

Demonstration : Take a small amount of copper hydroxide in a test tube and add carefully a solution of hydrochloric acid till the precipitate of copper hydroxide dissolves to give a blue coloured solution. Evaporate a little of this solution to dryness.

A residue of copper chloride ($CuCl_2$) is obtained.

$2HCl$	+	$Cu(OH)_2$	=	$CuCl_2$	+	$2H_2O$
Hydrochloric acid (solution)		Copper hydroxide (blue)		Copper chloride (blue) (solution)		

Reactions between acids and bases yielding salt and water are called *reactions of neutralization.* Other acids and bases also take part in reactions of neutralization.

From an observation of the reactions of neutralization we have come across so far, one could notice that in these reactions the *hydrogen of the acids and the hydroxyl (OH) group of the bases combine together* to form molecules of water, and metal of the bases and the acid radicals combine together to form the salt.

Here is another example :

$H \cdot NO_3$	+	$K : OH$	=	KNO_3	+	H_2O
(acid)		(base)		(salt)		(water)

(*e*) *Interaction between acids and salts—reaction of double decomposition.*

Let us now study what happens when a salt reacts with an acid.

Experiment : Take a little sodium chloride in a dry test tube and pour carefully about one ml of strong sulphuric acid into it. Warm the test tube and see if any strong fumes are given off. Dip a glass rod in a bottle of ammonia and bring the moist tip near the mouth of the test tube. What do you see ?

Wash the rod, dip it again in ammonia and bring the tip near the mouth of the bottle of concentrated hydrochloric acid. Do you see the similarity in the two cases ?

This reaction is explained by the following equation :

H_2SO_4	$+$	$2NaCl$	$=$	Na_2SO_4	$+$	$2HCl$
Sulphuric acid (acid)		Sodium chloride (salt)		Sodium sulphate (another salt)		Hydro-chloric acid (another acid)

Experiment : Take a small quantity of calcium carbonate in a test tube and pour 2 ml of water in it. Does it dissolve ? Then carefully add 2 ml dilute hydrochloric acid and observe what happens Bring a lighted splinter into the mouth of the test tube without touching the sides. What happens ? Hold a glass tube with a drop of lime water near the mouth of the test tube. What happens to the lime water ?

What is left in the test tube ? Is it a clear solution ? Evaporate a few drops of this solution. What conclusion can you draw ?

In this case the insoluble calcium carbonate has dissolved in hydrochloric acid to give a new salt calcium chloride and carbonic acid (a new acid). This decomposes into carbon dioxide gas which is the anhydride of carbonic acid and water.

$$\underset{\text{Hydrochloric acid}}{2HCl} + \underset{\text{Calcium carbonate}}{CaCO_3} = \underset{\text{Calcium chloride (new salt)}}{CaCl_2} + \underset{\text{carbonic acid (new acid)}}{H_2CO_3}$$

$$H_2CO_3 \downarrow H_2O + \underset{\text{carbon dioxide (carbonic anhydride}}{CO_2}$$

Similarly by action of sulphuric acid on barium nitrate, nitric acid and an insoluble precipitate of barium sulphate are obtained as follows :

$$\underset{\text{Sulphuric acid (solution)}}{H_2 : SO_4} + \underset{\text{Barium nitrate (salt in solution)}}{Ba : (NO_3)_2} = \underset{\text{Barium sulphate (insoluble precipitate)}}{BaSO_4 \downarrow} + \underset{\text{Nitric acid (in solution)}}{2HNO_3}$$

The above mentioned reactions of acids, with oxides, bases and salts, are examples of interaction between two compounds, as a result of which two new compounds are obtained. Such chemical reactions are called *reactions of double decomposition.* In such reactions molecules of two compounds interchange their parts. The interaction between sulphuric acid and barium hydroxide is another example of double decomposition where interchange of atoms of hydrogen and barium in the molecules of the acid and the base takes place. At the same time, hydroxyl groups (OH) are interchanged with the acid radical ($=SO_4$) of the acid.

(f) Basicity of acids—composition of acids— oxy and non-oxy acids

Study of the interaction of acids with metals, oxides, bases and salts has shown that in these reactions the hydrogen atoms of the acid molecules are displaced by the atoms of metals present in the molecules of metal oxides, bases and salts. The number of hydrogen atoms that can be displaced from the molecules of different acids is not the same.

See the formula of the acids from the following table.

Acid	*Number of hydrogen atoms*	*Acid radical*	*Basicity*
HCl	1	$(-Cl)$	Monobasic
HI	1	$(-I)$	,,
HNO_3	1	$(-NO_3)$	,,
H_2S	2	$(=S)$	Dibasic
H_2SO_4	2	$(=SO_4)$	,,
H_2CO_3	2	$(=CO_3)$	,,
H_3PO_4	3	$(\equiv PO_4)$	Tribasic

From this table it is clear that in the case of HCl, HI and HNO_3 only one hydrogen atom can be displaced by a metal and these acids are called *monobasic*. Similarly in H_2S, H_2SO_4 and H_2CO_3 it is possible to displace two atoms of hydrogen from each molecule and these are called *dibasic* acids. On the same basis H_3PO_4 is known as a *tribasic* acid.

The basicity of an acid indicates the maximum number of hydrogen atoms in a molecule of the acid that could be displaced by or interchanged with metal atoms.

It is therefore numerically equal to the valency of the acid radical (see table).

In molecules of sulphuric, nitric and phosphoric acids, atoms of hydrogen are joined to remnants (radicals) of acids which consist of atoms of non-metals and oxygen. These acids are called *oxyacids* since all of them contain oxygen. Besides these oxyacids there are many other acids. In a molecule of hydrochloric acid, HCl, however an atom of hydrogen is connected to chlorine. The radical of this acid is the chloride radical. It does not contain any atom of oxygen in it and therefore hydrochloric acid is a *non-oxyacid*. Hydrogen sulphide, H_2S, also belongs to this group of acids. It is a solution of

gaseous hydrogen sulphide in water. The radical of this acid is the *sulphide radical.*

Thus molecules of all acids (both oxyacids and non-oxyacids) consist of atom or atoms of hydrogen connected to the remnants (radicals) of the acids. Acids may be defined anew as follows :

Acids are compounds whose molecules contain atoms of hydrogen which could be displaced by atoms of metals or interchanged with them.

Balancing of Equations

In the equation of a reaction of neutralization it is necessary to write the formulae of substances and to equalize the number of atoms of the various elements on the left and right sides of the equation. This is in accordance with the law of conservation of mass.

For instance, in the reaction of neutralization of phosphoric acid with calcium hydroxide, the salt obtained is calcium phosphate which is represented by the formula $Ca_3(PO_4)_2$, since calcium is bivalent, whereas phosphate radical, (PO_4) is trivalent. The reaction may be written as :

$$Ca(OH)_2 + H_3PO_4 \rightarrow Ca_3(PO_4)_2 + H_2O$$

In the above expression the number of atoms of the same element are not equal on both sides of the equation. (An equation in which the above condition is satisfied is known as a balanced equation). In a molecule of calcium phosphate there are three atoms of calcium and two phosphate radicals. Since each acid molecule has only one phosphate radical and each molecule of the base has only one atom of Ca, a molecule of $Ca_3(PO_4)_2$ would be formed from 3 molecules of the base and two molecules of the acid. Therefore,

$$3Ca(OH)_2 + 2H_3PO_4 \rightarrow Ca_3(PO_4)_2 + H_2O.$$

The number of molecules of water on the right-hand side of the equation must be equal to the number of hydroxyl groups which form water by combination with hydrogen atoms of acids. As 3 molecules of calcium hydroxide contain 6 hydroxyl groups, they will form 6 molecules of water, since each OH group will give rise to a molecule of water. It is therefore necessary to put the coefficient 6 before the

formula of water. The final equation of the reaction can then be represented as :

$3Ca(OH)_2$	+	$2H_3PO_4$	=	$Ca_3(PO_4)_2$	+	$6H_2O$
Base (Calcium hydroxide)		Acid (Phosphoric acid)		Salt (Calcium Phosphate)		Water

We have now the balanced equation of the reaction.

Uses Of Acids

Acids find many uses in various industrial processes. These are very important in the development of national economy. Sulphuric acid is the most important amongst these acids. At present, the largest amount of sulphuric acid is used in the production of mineral fertilizers. Sulphuric acid is used for preparing many other acids and salts. The other areas where it finds use are in the manufacture of rayon and staple fibres, alum, petrol refining explosives, drugs and dyes. Nitric acid is used in the production of mineral fertilizers. This acid is also employed in the production of artificial fibres, dyes and explosives. Hydrochloric acid is used in the manufacture of different salts.

Questions

1. You are provided with 3 colourless liquids in 3 test tubes. One of them is acidic, another is alkaline and the third is neutral. How will you identify each ?
2. Write the equations of the following reactions :
   (*a*) Calcium oxide with sulphuric acid,
   (*b*) Iron oxide with sulphuric acid,
   (*c*) Zinc oxide with hydrochloric acid,
   (*d*) Copper oxide with nitric acid,
   (*e*) Ferric oxide with nitric acid.
3 Write the equations of the reactions of the following hydroxides with sulphuric acid :
   (*a*) Magnesium hydroxide.
   (*b*) Copper hydroxide.
   (*c*) Ferric hydroxide.
   To what type do these reactions belong ?
   Name the salts which are obtained.

4. What are the chemical properties of acids ? Write the equations of the reactions that characterize these properties.

5. What is a 'reaction of double decomposition' ? Give two examples.

6. How can the following salts be obtained by the interaction of oxides of metals with acids ? Write the equations of the reactions involved.

   (*a*) Zinc nitrate.

   (*b*) Zinc sulphate.

   (*c*) Zinc phosphate.

7. Write the equation of the reactions of hydrochloric acid and sulphuric acid on aluminium.

8. What is the "basicity" of an acid ? Classify the following into monobasic, dibasic and tribasic acids :

   HCl, HNO_3, H_2SO_4, HPO_3, H_3PO_4, H_2CrO_4, $HClO_3$, H_2CO_3.

9. Give examples of substances which would give salts by the action of acids. Which of them are simple substances and which of them are compounds ? Classify the following substances on the above basis :

   Zn, $NaOH$, CaO, Mg, KOH, Fe_2O_3, $Cu(OH)_2$, Fe, Cu.

10. How may the following acids be classified according to their composition ?

    HBr, H_2CO_3, HI, H_2CrO_4, $HClO_3$, HNO_3.

11. Which chemical element is present in all acids ?

12. What is an acid radical ?

13. Why is sulphuric acid regarded as the most important acid ?

14. What are the uses of nitric and hydrochloric acids in industrial manufacture ?

Precaution for the use of acids and alkalis

Acids and alkalis are corrosive to skin, clothes etc. It is necessary to use them very carefully. The following precautions must be observed :

1. Solid alkali should never be picked up by hand. Tongs, pincers or spatula should invariably be used.
2. Any solid alkali that is left after use should be put into a container.
3. Concentrated acids especially sulphuric acid, should be diluted by pouring them into water *and not vice-versa.*
4. Place the stoppers of bottles containing acids or alkalies, upside down, on a glass plate. Do not place them on the surface of the table. Use only cork, rubber or plastic stoppers (and not glass stoppers) for alkali bottles.
5. Solutions of acids or alkalies should be poured from bottles into test tubes or beakers, carefully. Be careful to pour the last drop from the container into the vessel in which the liquid is being poured (see Fig. 9).

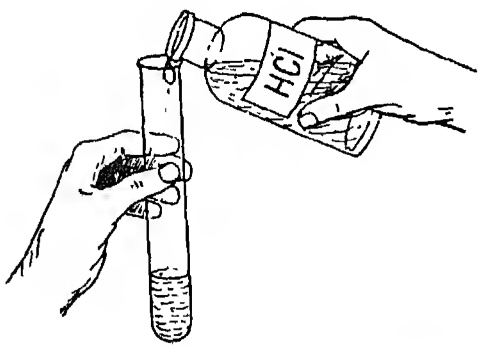


FIG. 9. Correct method of pouring reagents

6. While heating a solution of an acid or an alkali in a test tube, keep the mouth of the tube away from yourself and your neighbours.

There is a possibility of the boiling liquid spurting out of the test tube.

First aid in case of acid or alkali burns

If an acid or alkali falls on the skin, wash the place with running water for a couple of minutes. Thereafter put cotton wool moistened with a 3 percent solution of potassium permanganate on the affected part and bandage it. If some solutions of acid or alkali spurt into your eyes, wash with running water. Then go to a doctor immediately. In every case inform your teacher and take his advice.

Practical Work No. 2

Reaction of Neutralization

Equipment and reagents : Spirit lamp, burette with stand, pipette, beakers, porcelain dish, funnel, asbestos-covered wire gauze, a pair of tongs, a glazed tile (a white piece of paper may be used instead) and a dropper. Solutions of hydrochloric acid, sodium hydroxide, methyl orange and phenolphthalein.

Procedure : Clamp the burette to a stand vertically. Take the given solution of acid in a beaker and pour it into the burette with the help of a funnel. Drain off some portion of the acid into the beaker so as to remove the air from the jet of the burette. This ensures that the entire space below is filled with the liquid. Adjust the level of the liquid to any mark in the burette and note the reading in your note book. Keep your eye at the level of the liquid in the burette while taking the reading. With the help of a pipette take a known volume of alkali in a clean beaker.

Place the beaker on a white sheet of paper under the burette in such a way that its nozzle just enters the mouth of the beaker. Add two drops of methyl orange into the beaker with a dropper or a glass tube and note the colour. Run down small quantities of the acid from the burette starting from one to two ml and coming down to drop by drop. After every addition of the acid, shake the beaker carefully with a circular motion. Note the appearance of a change in

colour. This shows that the alkali in the beaker has been completely neutralized. Note the level of the acid now in the burette. Calculate how many millilitres of acid were spent to neutralize the given alkali solution.

Take some drops of the solution from the beaker, and evaporate to dryness on a watch glass. What do you observe ?

Write a report of the work done.

Phenolphthalein can also be used as the indicator in place of methyl orange. In that case, the whole of the alkali is neutralized by the acid when the pink colour of the solution just disappears and the solution becomes colourless.

Note: After your work is finished, clean the apparatus and your working place.

Questions

1. What are the characteristic chemical properties of acidic and basic oxides ? Give the equations of the reactions that represent them.
2. Calculate the percentage of sulphur in (*a*) sulphurous anhydride and (*b*) sulphuric anhydride. (Ans. (*a*) 50%; (*b*) 40%)
3. Which of these substances—P_2O_5, CaO, HNO_3, $Ba(OH)_2$, MgO, SO_3—will react with sulphuric anhydride ? Write the equations of the reactions.
4. Give in a tabular form, as shown below, the composition and properties of the oxides studied:

Basic oxides					*Acidic oxides*				
Name	Formula	Action of			Name	Formula	Action of		
		acids	acidic oxide	water			base	basic oxide	water

BASES

You have already learnt that bases are those compounds whose molecules generally consist of atoms of metals connected with hydroxyl groups. Examples of bases are : Sodium hydroxide (Caustic Soda)—$NaOH$, Calcium hydroxide (Slaked lime)—$Ca(OH)_2$, Ferric hydroxide—$Fe(OH)_3$ and Copper hydroxide—$Cu(OH)_2$. Bases can be divided into two groups : soluble bases (also called alkalies)and insoluble bases. Bases interact in the 'reaction of neutralization' with acids to form salt and water Hence a more complete definition, to show the composition of the molecules of bases and their characteristic property can be as follows :

Bases are compounds whose molecules generally consist of one or more hydroxyl groups attached to a metal atom and which by interaction with acids give salt and water

Bases have the general formula $M(OH)_n$, where 'M' stands for the atom of a metal and 'n' is the number of hydroxyl groups in a molecule of the base. The number of hydroxyl groups present in the molecule of a base (n) is equal to the valency of the metal present.

6. Chemical Properties of Bases.

(*a*) *Action of solutions of alkali on indicators*

As you already know, one of the characteristic properties of alkalies is their ability to change the colour of indicators. In the laboratory, methyl orange is often used, besides litmus and phenolphthalein as an indicator. The orange colour of the solution of this indicator changes into yellow in the presence of an alkali.

(*b*) *Action of heat on bases*

Demonstration : Take a few granules, flakes or pieces of solid caustic soda by means of a spatula in a dry test tube. Warm carefully and then heat strongly. See that it melts but there is no evidence of a change taking place. Allow it to cool and examine its appearance. Dissolve the residue in water and test with litmus.

Soluble bases (alkalies)are stable to heat. Caustic soda melts on heating but does not change chemically, even at the temperature of

boiling (1388°C). Caustic potash is also stable. Calcium hydroxide begins to decompose into water and calcium oxide at 450°C.

Demonstration : Take 2-3 g of copper hydroxide and ferric hydroxide in separate test tubes. Show the students that these insoluble bases have distinct colours. Heat the test tubes by turn and show the change in colour of the residue and formation of water droplets on the sides of the test tubes.

Copper hydroxide decomposes into water and copper oxide even on mild heating.

$$\underset{\text{Copper hydroxide}}{Cu(OH)_2} \xrightarrow{\text{heat}} \underset{\text{Copper oxide}}{CuO} + \underset{\text{Water}}{H_2O}$$

Iron hydroxide also decomposes easily into water and iron oxide (ferric) :

$$\underset{\text{Ferric hydroxide}}{2Fe(OH)_3} \xrightarrow{\text{heat}} \underset{\text{Ferric oxide}}{Fe_2O_3} + \underset{\text{Water}}{3H_2O}$$

Many other insoluble bases also decompose in the same manner.

Reactions of decomposition in which one of the products is water are known as dehydrations.

(*c*) *Interaction of bases with acids*

You have already learnt that the ability to take part in the reaction of neutralization is a characteristic property of both acids and bases. Both soluble and insoluble bases take part in this reaction.

Examples of the reactions between bases and acids have been described in section 5.

You have already seen how carbon dioxide (anhydride of carbonic acid) reacts with caustic soda, forming sodium carbonate and water.

$$\underset{\text{Sodium hydroxide}}{2NaOH} + \underset{\text{Carbon dioxide}}{CO_2} = \underset{\text{Sodium carbonate}}{Na_2CO_3} + \underset{\text{Water}}{H_2O}$$

Salt and water are also obtained by the interaction of carbonic anhydride (carbon dioxide) with a solution of calcium hydroxide (lime water). The salt obtained in this reaction is calcium carbonate :

$Ca(OH)_2$	+	CO_2	=	$CaCO_3$	+	H_2O
Base		Anhydride		Salt		Water
(Calcium hydroxide)		(carbonic)		(Calcium carbonate)		

This reaction explains the formation of turbidity in lime water by passing carbon dioxide through it due to the formation of insoluble calcium carbonate.

The reaction with anhydrides of acids is a characteristic property of soluble bases (alkalies).

(*d*) *Interaction of soluble bases with salts*

Experiment : Take three test tubes A, B and C. Pour 2-3 ml of the following salt solutions in the three test tubes respectively—copper sulphate, ferric chloride and sodium sulphate.

Add by a dropper 10-15 drops of caustic soda into the test tube A. Observe if any change takes place. Shake the test tube and again observe.

Repeat the same operation with test tube B.

Pour 10 drops of barium hydroxide solution into the tube C similarly and observe the change that takes place.

The above experiments show that the interaction of soluble bases with solutions of salts gives insoluble residues or precipitates. Thus by mixing together solutions of caustic soda and copper sulphate a blue precipitate of copper hydroxide is obtained :

$2NaOH$	+	$Cu(SO)_4$	=	$Cu(OH)_2\downarrow$	+	Na_2SO_4
Base		Salt		Base		Salt
(Sodium hydroxide) (in solution)		(Copper sulphate (in solution)		(Copper hydroxide) (precipitate)		(Sodium sulphate (in solution)

Again, by the interaction of ferric chloride with caustic soda, ferric hydroxide is precipitated :

$3NaOH$	+	$FeCl_3$	=	$Fe(OH)_3 \downarrow$	+	$3NaCl$
Base		Salt		Base		Salt
(Caustic soda)		(Ferric chloride		(Ferric hydroxide)		(Sodium chloride)
(in solution)		(in solution)		(precipitate)		(in solution)

In the two cases mentioned above, reaction between solutions of salts and alkalies forms bases that are insoluble in water. Reaction between barium hydroxide and sodium sulphate gives the new base, sodium hydroxide, which is soluble. But in this case, however, it is the salt that is precipitated :

$Ba(OH)_2$	+	Na_2SO_4	=	$2NaOH$	+	$BaSO_4 \downarrow$
Base		Salt		Base		Salt
(Barium hydroxide)		(Sodium sulphate)		(Sodium hydroxide)		(Barium sulphate)
(in solution)		(in solution)		(in solution)		(precipitate)

For determining whether precipitation will take place in a chemical reaction between solutions of alkalies and salts, it is necessary to know the solubility in water of the products of the reaction.

The table on page 33 gives the solubility of salts and bases in water.

The use of this table can be illustrated as follows.

Suppose it is necessary to obtain magnesium hydroxide, $Mg(OH)_2$. Note from the table that this substance is sparingly soluble in water. Therefore, it is possible to obtain it by mixing solutions of any soluble salt of magnesium and of an alkali. From the table you can find that magnesium chloride, magnesium sulphate and magnesium nitrate are the soluble salts, any one of which could give magnesium hydroxide. Sodium hydroxide may be taken as the alkali. For example :

$Mg(NO_3)_2$	+	$2NaOH$	=	$Mg(OH)_2 \downarrow$	+	$2NaNO_3$
Salt		Base		Base		Salt
(Magnesium nitrate)		(Sodium hydroxide)		(Magnesium hydroxide)		(Sodium nitrate)
(in solution)		(in solution)		(precipitate)		(in solution)

SOLUBILITY OF SALTS AND BASES IN WATER

	K I	Na I	Ba II	Ca II	Mg II	Al III	Cr III	Fe II	Fe III	Mn II	Zn II	Ag I	Hg II	Pb II	Sn II	Cu II
OH(I)	S	S	S	Sp	Sp	N	N	N	N	N	N	—	—	N	N	N
Cl(I)	S	S	S	S	S	S	S	S	S	S	S	N	S	S	S	S
S(II)	S	S	S	Sp	S	—	—	N	—	N	N	N	N	N	N	N
SO_3(II)	S	S	N	N	N	—	—	N	—	N	N	N	N	N	—	N
SO_4(II)	S	S	N	Sp	S	S	S	S	S	S	S	Sp	S	N	S	S
PO_4(III)	S	S	N	N	N	N	N	N	N	N	N	N	N	N	N	N
CO_3 II)	S	S	N	N	N	—	—	N	N	N	N	N	N	N	—	N
SiO_3(II)	S	S	N	N	N	N	N	N	N	N	N	N	—	N	—	N
NO_3(I)	S	S	S	S	S	S	S	S	S	S	S	S	S	S	—	S
CH_3COO(I)	S	S	S	S	S	Sp	S	S	S	S	S	S	S	S	S	S

Legend

S Soluble

N Insoluble

— Does not exist

Sp Sparingly soluble

Roman numerals indicate the valency of the radical or metal.

Uses Of Bases

Many bases are very important in industry. Sodium hydroxide, also known as caustic soda, is especially important. It is used in the manufacture of soaps. It is also used in production of artificial silk, organic dyes, paper, in the textile industry, in the purification of oil products, in the production of aluminium oxide and also in many other branches of industry.

Other bases are also used in industry in large quantities. For example, calcium hydroxide or slaked lime is used for the preparation of mortar and for white washing the buildings.

Questions

1. How can you distinguish a solution of an alkali from a solution of an acid ?
2. Write the formulae of bases known to you. Write the equations of the reactions of interaction of these bases with nitric and sulphuric acids.
3. What are the common properties of bases ? What is the difference between soluble and insoluble bases ?
4. Chromium, zinc, aluminium and magnesium hydroxides on heating give the following oxides respectively :

   Cr_2O_3, ZnO, Al_2O_3, MgO.

   Write the equations of the reactions of dehydration of the above hydroxides.
5. To what type of reaction does the interaction of solutions of an alkali and a salt belong ?
6. How can you get insoluble bases ? Do you find any relation between solubility of a base and its stability to heat ?
7. What are the substances that give salts by interaction with bases ?
8. What are the common properties of bases ? What is the difference between soluble and insoluble bases ?

Practical Work No. 3

Preparation of insoluble bases by interaction of solutions of alkalis and salts.

Apparatus : Test tubes, beakers, metallic stand with ring, funnels, wash bottle, spirit lamp, glass rod, filter paper.

Reagents : Solutions of copper nitrate, magnesium chloride and caustic soda.

Procedure : Pour into a beaker 20-25 ml of the caustic soda solution and add to it the solution of copper nitrate in small portions of about 2 ml. stir and let it stand. The solution above the precipitate should be colourless. (What will you do if the solution is of blue colour ?) Decant carefully the colourless solution into a funnel fitted with a filter paper. The filter should be prepared while the solution is kept standing. Add some water from the wash bottle (see Fig. 10) into the beaker containing the precipitate and mix the contents with a glass rod. (Be careful not to break the beaker). Pour the contents on the filter and wash the precipitate on the filter with water from the wash bottle. Record your experiment and note the properties of the copper hydroxide obtained.

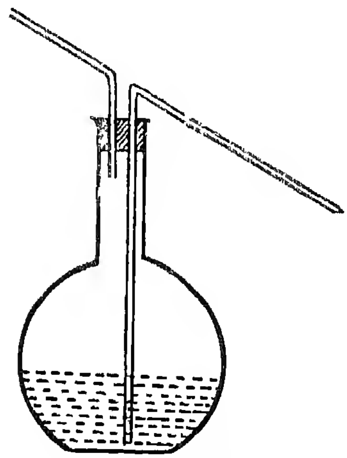
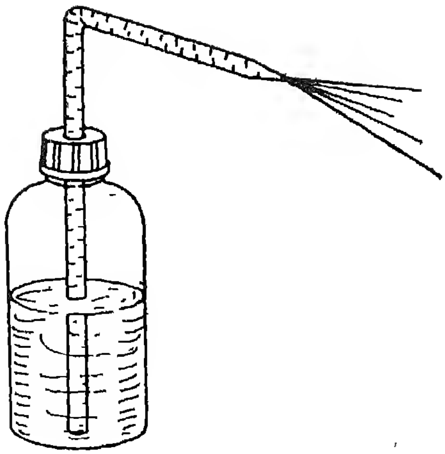


FIG. 10 (a) One type of wash bottle

Prepare magnesium hydroxide by the same method. For this experiment take 20 ml of a solution of caustic soda and 20 ml of a solution of magnesium chloride. Make a report of your work.

How can you prove that the substances obtained in both cases are bases ?

Clean the glassware and the working place after finishing your work.



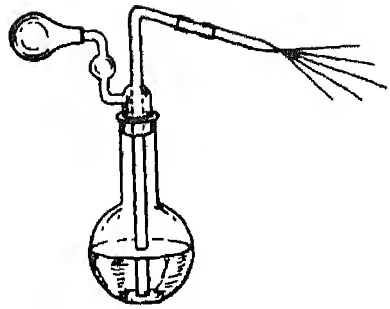


FIG 10 (b) Types of wash bottles FIG. 10 (c)

7. Calculations Using Equations of Chemical Reactions

With the help of equations of chemical reactions we can calculate both the quantity of the initial substances to be taken for the reaction and the quantity of products obtained. For example let us write the equation of reaction between magnesium oxide and sulphuric acid and calculate the molecular weights of the initial substances and the products :

	MgO	$+$	H_2SO_4	$=$	$MgSO_4$	$+$	H_2O
Mol. Wt.	40		98		120		18

From the law of conservation of mass it follows that interaction of 40 parts by weight of magnesium oxide with 98 parts by weight of sulphuric acid gives 120 parts by weight of salt (magnesium sulphate) and 18 parts by weight of water. Whatever be the unit of weight employed in the case of the initial substances (gram, kilogram, ton, etc.) the final products would also be obtained in the same units.

Example 1

How many grams of sulphuric acid would 100 grams of MgO require for complete reaction ?

Method : Write down the equation of the reaction and the molecular weights of the reactants below their respective formulae. The actual weights taken are written above, as shown below :

$$\underset{40}{\overset{100\text{ g}}{MgO}} + \underset{98}{\overset{(?\text{ g})}{H_2SO_4}} = MgSO_4 + H_2O$$

From the equation we know that :

From 40 g of MgO the amount of $MgSO_4$ obtained is

$\therefore$ 1 ,, ,, ,, ,, ,, ,, $\frac{98}{40}$,, ,, ,,

$\therefore$ 100 ,, ,, ,, ,, ,, ,, $\frac{98}{40} \times 100$ g ,,

$= 245$ g

Answer : For interaction with 100 gm of MgO, 245 g of sulphuric acid would be required.

Example 2

How many grams of magnesium sulphate would be obtained by the action of sulphuric acid on 16 g of MgO ?

Method : Write the equation of the reaction and the molecular weights of magnesium oxide and magnesium sulphate below their respective formulae and the actual weights taken above, as shown below :

$$\underset{40}{\overset{16\text{ g}}{MgO}} + H_2SO_4 = \underset{120}{\overset{(?\text{ g})}{MgSO_4}} + H_2O$$

From the equation we know that :

From 40 g of MgO the amount of $MgSO_4$ obtained is

$= 120$ g

$\therefore$,, 1 ,, ,, ,, ,, ,, ,, $= \frac{120}{40}$ g

$\therefore$,, 16 ,, ,, ,, ,, ,, ,, $= \frac{120}{40} \times 16$ g

$= 48$ g

Answer : By the action of sulphuric acid on 16 g of MgO, 48 g of $MgSO_4$ would be obtained.

Thus, there is a proportional dependence of the final products on the initial substances .

Hence the above problem may also be written as a proportion, as $\frac{40}{120}=\frac{16}{x}$ where x stands for the unknown quantity to be calculated.

From the proportion it follows that $x=\frac{16\times 120}{40}=48$ g

Answer : By the action of sulphuric acid on 16 g of MgO, 48 g of $MgSO_4$ is obtained.

Problems

1. How many grams of sulphuric acid are required for complete interaction with 48 g of ferric oxide ? (Ans. 88 2 g)
2. If nitric acid reacts completely with 224 g of CaO, how many grams of the salt would be obtained ? (Ans. 656 g)
3. How many grams of nitric acid are necessary for the complete neutralization of 20 g of caustic soda ? (Ans. 31 5 g)
4. How many grams of zinc chloride are obtained by the action of hydrochloric acid on 1.62 g of zinc oxide. (Ans. 2.72 g)
5. How many grams of barium sulphate are precipitated by the action of sulphuric acid on a solution containing 5 22 g of barium nitrate ? (Ans. 4.66 g)
6. How many grams of HNO_3 are required for reacting completely with 20 g of copper oxide ? (Ans. 31.5 g)
7. How many litres of hydrogen are evolved by the action of 130 g of zinc on a solution of sulphuric acid ? (1 litre of hydrogen at NTP weighs 0.089 g) (Ans. 44 8 litre)
8. How many litres of hydrogen could be obtained by the action of aluminium on a solution containing 98 g of sulphuric acid ? (Ans. 22.4 litres)
9. How many grams of ferric hydroxide are obtained by the interaction of 12 grams of caustic soda with a solution of ferric chloride? (Ans. 10.7 g)

10. How many grams of salt will be obtained by the interaction of 33.6 grams of caustic potash with a solution of sulphuric acid ? (Ans. 52.2 g)
11. How many grams of copper oxide will be obtained by heating 2.45 grams of copper hydroxide ? (Ans. 2 g)

SALTS

You have learnt that salts are compounds whose molecules contain atoms of metals connected to acid remnants (radicals).

Salts are produced by interaction of acids and bases. They may also be produced by interaction of metallic oxides with acids or by displacement of the atoms of hydrogen of acids by atoms of metals.

8. Chemical Properties of Salts

You have already come across some of the characteristic reactions of salts.

Interaction of salts with acids.

Task : Using the table of solubility, find out how barium sulphate can be obtained from sulphuric acid.

Perhaps you may come up with an equation as follows:

$$\underset{\text{Barium chloride}}{BaCl_2} + \underset{\text{Sulphuric acid}}{H_2SO_4} = \underset{\text{Barium sulphate}}{BaSO_4\downarrow} + \underset{\text{Hydrochloric acid}}{2HCl}$$

It is an example of the interaction of salts with acids. Let us now see some more examples.

Experiment : Add dilute hydrochloric acid to a solution of lead nitrate in a test tube. What do you observe ? Write the equation of the reaction.

The white precipitate obtained is lead chloride which may be expressed as follows :

$$\underset{\substack{\text{Lead nitrate}\\\text{(in solution)}}}{Pb(NO_3)_2} + \underset{\substack{\text{Hydro-chloric acid}\\\text{(in solution)}}}{2HCl} = \underset{\substack{\text{Lead chloride}\\\text{(precipitate)}}}{PbCl_2\downarrow} + \underset{\substack{\text{Nitric acid}\\\text{(in solution)}}}{2HNO_3}$$

In this reaction an insoluble salt is formed

Experiment : Take a solution of sodium carbonate in a test tube and add dilute sulphuric acid into it in small quantities at a time. What do you observe ?

In this reaction carbonic acid is produced which decomposes into its anhydride (carbon dioxide) and water.

The same reaction takes place when dilute acids interact with dry sodium carbonate :

$$\underset{\text{(in solution or dry)}}{Na_2CO_3} + \underset{\text{(in solution)}}{H_2SO_4} = \underset{\text{(in solution)}}{Na_2SO_4} + \underset{\text{(in solution)}}{H_2CO_3}$$

$$\underset{\text{(in solution)}}{H_2CO_3} = H_2O + \underset{\text{gas}}{CO_2\uparrow}$$

In this reaction a gaseous product (CO_2) has been formed.

Demonstration : Take some powdered potassium nitrate in a dry test tube and add a few drops of strong sulphuric acid carefully. Warm the test tube gently and observe oily drops collecting on the sides of the test tube and brown fumes coming out.

This is due to the formation of nitric acid as expressed below :

$$(i)\quad \underset{\substack{\text{Potassium}\\ \text{nitrate}}}{2KNO_3} + \underset{\substack{\text{Sulphuric}\\ \text{acid}}}{H_2SO_4} = \underset{\substack{\text{Potassium}\\ \text{sulphate}}}{K_2SO_4} + \underset{\substack{\text{Nitric}\\ \text{acid}}}{2HNO_3}$$

$$(ii)\quad \underset{\substack{\text{Nitric}\\ \text{acid}}}{4HNO_3} \xrightarrow{\text{heat}} 2H_2O + \underset{\substack{\text{Nitrogen}\\ \text{dioxide}\\ \text{(brown}\\ \text{fumes)}}}{4NO_2} + \underset{\text{Oxygen}}{O_2}$$

Take another dry test tube with potassium sulphate and pour carefully a few drops of strong nitric acid, warm this and observe.

In this case also only nitric acid is formed as is shown by the oily drops and brown fumes. The vapours are collected by distillation and are found to contain only nitric acid; no sulphuric acid is produced.

$$K_2SO_4 + 2HNO_3 \xrightarrow{\text{does not give}} KNO_3 + H_2SO_4$$

Nitric acid (b.p. 86°C) is more volatile than sulphuric acid (b.p 338°C). In such double decompositions only the more volatile acid is formed as a result of the interaction.

From the above experiments we can conclude that *the reaction of double decomposition between salts and acids takes place when*, as a result of the reaction,

(*i*) an insoluble salt is formed as a product ;

(*ii*) a gaseous or more volatile substance is formed as a product ;

(*iii*) an insoluble acid is produced.

9. Interaction of Salts With Bases

You are acquainted with the interaction of bases with salts. Perform the following experiment.

Experiment : Take two test tubes. Fill them with 3–4 ml of ferrous sulphate and sodium carbonate solutions, respectively. Add a solution of caustic soda in the first and lime water in the other separately till you find some definite change taking place in each.

In the first case a dirty green precipitate of ferrous hydroxide is produced and in the other a white precipitate of calcium carbonate is produced. The reactions are represented as follows :

$$\underset{\text{(solution)}}{FeSO_4} + \underset{\text{(solution)}}{2NaOH} = \underset{\text{(precipitate)}}{Fe(OH)_2} + \underset{\text{(solution)}}{Na_2SO_4}$$

$$\underset{\text{(solution)}}{Na_2CO_3} + \underset{\text{(solution)}}{Ca(OH)_2} = \underset{\text{(solution)}}{2NaOH} + \underset{\text{(precipitate)}}{CaCO_3\downarrow}$$

Task

Using the table of solubility, select two other examples of interaction of salts with bases so that:

(1) an insoluble base and a soluble salt are obtained as products ;

(2) a soluble base and an insoluble salt are obtained as products.

Write the equation of the reactions.

10. Interaction of Salts With one Another

What are the conditions necessary for the interaction of salts with one another? What are the indications of these reactions? Let us perform the following experiments to obtain answers to these questions.

Experiment 1: Add a solution of barium chloride to a solution of sodium sulphate in a test tube. What do you observe? Write the equation of the reaction. What is the substance precipitated? To what type of reactions does this interaction belong?

Experiment 2: Take a solution of silver nitrate in a test tube and add to it a solution of sodium chloride. What do you observe? Write the equation of the reaction.

Experiment 3: Mix together solution of sodium nitrate and calcium chloride in a test tube. Do you notice any indication of a chemical reaction?

Experiment 4: Take a small quantity of a solution of potassium sulphate in a test tube, and add to it a solution of sodium chloride. Do you observe any indication of a chemical reaction?

It may be observed, from the above reactions, that *in some reactions of double decomposition between two salts, two new salts are obtained, one of which is precipitated* as in experiments 1 and 2:

$$\underset{\text{(in solution)}}{BaCl_2} + \underset{\text{(in solution)}}{Na_2SO_4} = \underset{\text{(precipitate)}}{BaSO_4} \downarrow + \underset{\text{(in solution)}}{2NaCl}$$

$$\underset{\text{(in solution)}}{AgNO_3} + \underset{\text{(in solution)}}{NaCl} = \underset{\text{(precipitate)}}{AgCl} \downarrow + \underset{\text{(in solution)}}{NaNO_3}$$

At the same time it may be observed that *no precipitate is obtained when the reactants and possible products are all soluble* as in experiments 3 and 4.

11. Interaction of Salts with Metals. The Activity Series of Metals

In the reaction of displacement you have learnt that by the action of iron on a solution of copper chloride, copper is produced.

$$Fe + CuCl_2 = FeCl_2 + Cu \downarrow$$

Let us now study in more detail the interaction of other metals with solution of salts.

Experiment : Take 5 test tubes and number them. Pour into each of them up to a third of capacity the solutions of the following salts. Put two pieces of zinc into each of the first two test tubes, two pieces of lead into the third and two pieces of copper into each of the last two test tubes.

Make the following table in your note book to record your observations.

No. of test tube	*Salt solution contained*	*Metal added*	*Observation*
1.	$Pb(NO_3)_2$	Zn	
2.	$Cu(NO_3)_2$	Zn	
3.	$Cu(NO_3)_2$	Pb	
4.	$Pb(NO_3)_2$	Cu	
5	$Zn(NO_3)_2$	Cu	

In which test tube do you observe the appearance of a new substance ?

Write the equation of the reaction.

These experiments show that zinc displaces lead and copper from the solutions of their salts. So zinc is considered more active than lead and copper.

Lead displaces copper from solutions of its salts. So lead is considered more active than copper. But copper displaces neither lead nor zinc from solutions of their salts. So copper is less active than these metals.

Can copper displace any metal from a solution of a salt ?

Experiment : Take a solution of mercuric nitrate and dip a copper plate in it. Observe a shining layer of mercury on the surface of the copper plate after some time.

Copper displaces mercury from the solution of its salt.

$$Cu + Hg(NO_3)_2 = Cu(NO_3)_2 + Hg \downarrow$$

If the plate is kept for a long time in the solution, the solution gets a blue colour. Therefore, there must be sufficient salt of copper present in the solution to give it a blue colour. Also, the weight of the copper plate becomes less. Thus copper is more active than mercury.

Many such experiments with different metals and salts have shown that activity differs from metal to metal. For example, iron is more active than copper. It displaces copper from the solutions of its salts. Copper is more active than mercury and so it displaces mercury from the solutions of its salts. But copper is less active than iron and cannot displace iron from the solutions of its salts.

Metals are arranged in the order of decreasing activity in a special row according to their activity :

K, Na, Ba, Mg, Al, Zn, Fe, Ni, Sn, Pb, (H), Cu, Hg, Ag, Pt, Au.

This row is called the *activity series* of metals. Hydrogen also finds a place in this series, even though it is not a metal.

If the place of a metal in the activity series is known, it can be predicted whether it will displace other metals from their salt solutions or hydrogen from acids. This is because all metals, which appear before hydrogen in the series (i.e. to the left of it) displace it from acid solutions (except from nitric acid). Further, each metal displaces other metals on its right from the solutions of their salts.

For example, zinc displaces all the metals on its right starting from iron, from solutions of their salts and hydrogen from acid solutions (except from nitric acid). Similarly copper displaces silver, mercury and gold from solutions of their salts but does not displace hydrogen from acids. It may be pointed out that the activity series of metals could only be used for the *reaction of displacement between metals and solutions of salts or acids*. This series does not explain the reaction of double decomposition, since 'free' metals do not take part in such reactions.

Reactions of displacement of metals from the solutions of their salts have considerable industrial importance.

Questions

1. Write the equations of the reactions by which the following conversions are effected :
   (*a*) $Zn \rightarrow ZnSO_4 \rightarrow Zn(OH)_2$
   (*b*) $Cu \rightarrow CuCl_2 \rightarrow Cu$
   (*c*) $Cu \rightarrow Cu(NO_3)_2 \rightarrow Cu(OH)_2 \rightarrow CuO \rightarrow Cu$
2. How much iron will displace 960 kg of copper from a solution of copper sulphate ? (**Ans.** 842 kg)
3. How many grams of caustic soda are necessary to obtain 32.1 g of ferric hydroxide from ferric nitrate ? (**Ans.** 36 g)
4. A solution of blue vitriol which was prepared for spraying over plants is kept into a bucket made of zinc. Why does the bucket leak after some time ?
5. How can you obtain copper from a solution of copper sulphate ?
6. If a solution of K_2SO_4 is added to a solution of 62.4 g of barium chloride, how many grams of $BaSO_4$ will be precipitated ? (**Ans.** 69.9 g)
7. How many litres of carbon dioxide are obtained by the action of hydrochloric acid on 80 g of marble ($CaCO_3$) ? (Weight of 1 litre of carbon dioxide at N.T.P. is 1.96 g) (**Ans.** 17.9 litres)
8. Write the equations of the following reactions :
   (*a*) Barium carbonate with nitric acid.
   (*b*) Copper sulphate with barium chloride.
9. Give one example each of reactions by which
   (*a*) an insoluble base and a soluble salt,
   (*b*) a soluble base and an insoluble salt
   can be obtained. Write equations.
10. Can we transport solutions of ferrous sulphate in copper or zinc containers ? Explain why ?
11. Starting with a zinc salt, how can we obtain zinc nitrate by the reaction of double decomposition ? Write the equation of the reaction.

12. How can you obtain copper nitrate from copper chloride by the reaction of double decomposition? Write the equation of the reaction.
13. Write two reactions of double decomposition by which zinc sulphate may be obtained.
14. How will you prepare magnesium nitrate from magnesium chloride? Give equations to represent the reactions you make use of.
15. How can you obtain caustic potash by the reaction of double decomposition?

12. Composition and names of salts

Salts may be considered as a product of displacement of the atoms of hydrogen of acids by the atoms of metals.

If the molecule of an acid contains two or more hydrogen atoms (polybasic acids) like H_2SO_4, it is possible to obtain compounds where one or both hydrogen atoms are displaced by metals. For example,

$$H_2SO_4 \nearrow Na_2SO_4 \text{ (Sodium sulphate)}$$
$$H_2SO_4 \searrow NaHSO_4$$

Sodium bisulphate or Sodium hydrogen sulphate.

The formation of these compounds, both of which are salts, may be expressed as follows:

$$\underset{\text{Acid}}{H_2SO_4} + \underset{\text{Base}}{2NaOH} = \underset{\substack{\text{Sodium sulphate} \\ \text{(normal salt)}}}{Na_2SO_4} + 2\,H_2O$$

$$\underset{\text{Acid}}{H_2SO_4} + \underset{\text{Base}}{NaOH} = \underset{\substack{\text{Sodium bisulphate} \\ \text{(acid salt)}}}{NaHSO_4} + H_2O$$

When all the atoms of displaceable hydrogen which form the molecule of an acid are displaced by atoms of metals, a normal salt is formed.

When only some atoms of displaceable hydrogen from the molecule of an acid are displaced by atoms of metals, an acid salt is obtained.

The acid remnant in the case of $NaHSO_4$ is the monovalent remnant $-HSO_4$.

Since only polybasic acids can give more than one acid remnant, acid salts are obtained only from these acids. For example, one can get two remnants $-HSO_4$ and $=SO_4$, corresponding to the acid salt and the normal salt, respectively, from sulphuric acid Monobasic acids do not give acid salts. As molecules of salts generally consist of atoms of metals and acid remnants (radicals) these salts are named by naming the metal first, and then the acid radical. This is illustrated in the following examples.

Formula	*Name of salt*	*Derived from acid*
$Ca(NO_3)_2$	Calcium nitrate	HNO_3
$NaCl$	Sodium chloride	HCl
$CuCO_3$	Copper carbonate	H_2CO_3
$NaHCO_3$	Sodium bicarbonate (acid)	H_2CO_3
$Al_2(SO_4)_3$	Aluminium sulphate	H_2SO_4
Na_3PO_4	Sodium phosphate (normal)	H_3PO_4
Na_2HPO_4	Disodium hydrogen phosphate (acid)	H_3PO_4
$BaCl_2$	Barium chloride	HCl
$AlBr_3$	Aluminium bromide	HBr
ZnS	Zinc sulphide	H_2S

If a metal gives rise to two salts having the same acid radical they are named as shown in the following examples :

(1) $FeSO_4$ — Ferr*ous* sulphate (valency of iron-2)
$Fe_2(SO_4)_3$ — Ferr*ic* sulphate (valency of iron-3)

(2) $FeCl_2$ — Ferr*ous* chloride (valency of iron-2)
$FeCl_3$ — Ferr*ic* chloride (valency of iron-3)

The name of the salt in which the metal exhibits lower valency will have the suffix "ous" while the name of salt in which it exhibits a higher valency will have the suffix "ic". This rule applies to the salts obtained from both oxyacids and non-oxyacids.

Sometimes common names instead of their chemical names are employed for substances

Formula	*Chemical name*	*Common name*
$NaCl$	Sodium chloride	Common salt
Na_2CO_3	Sodium carbonate	Common soda *or* washing soda
K_2CO_3	Potassium carbonate	Potash
$CuSO_4$	Copper sulphate	Blue vitriol
$AgNO_3$	Silver nitrate	Lunar caustic

Questions

(1) Write the formulae and names of the salts derived from the following acids with potassium and zinc :
HNO_2, H_2SO_3, HI, H_2CO_3

(2) How will you name the following salts represented by the formulae :
K_2S, $Cr_2(SO_4)_3$, $CrCl_3$,
$CaCO_3$, $Ca(HCO_3)_2$.

(3) Write the formulae of :
Cuprous chloride, Cupric chl
Ferric sulphate, Ferrous [illegible]de.

13. Inter-relationship of Inorganic Compounds

Previously substances were divided into simple substances and compounds. Now you have learnt about some important classes of

inorganic compounds. These substances may be classified according to the diagrammatic scheme (Fig. 11).

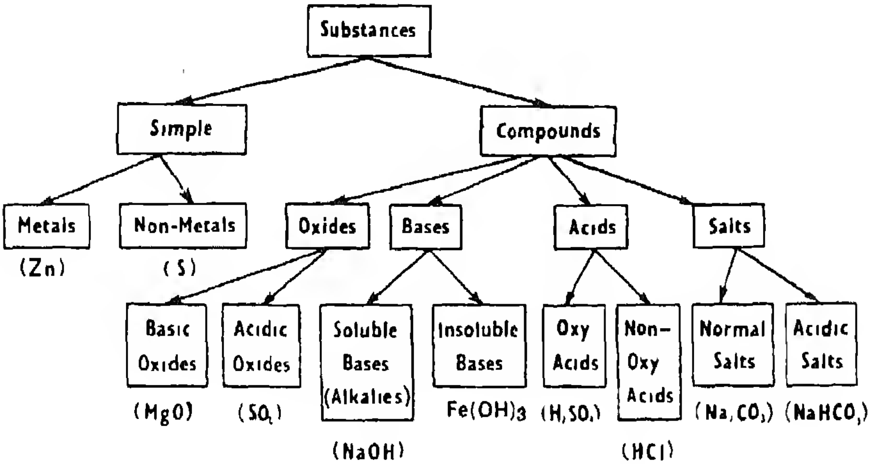


FIG. 11 Relationship of inorganic substances

(An example of each class is given in the brackets).

How are these classes of compounds related to each other ?

Let us take the simple substance, magnesium. By heating it in air or oxygen, we get the oxide, magnesium oxide.

$$2Mg + O_2 = 2MgO$$

This on hydration will give us the base magnesium hydroxide.

$$MgO + H_2O = Mg(OH)_2$$

Magnesium hydroxide will react with anhydrides of acids (e.g. carbonic anhydride) to form salts. Thus, in this case, the salt magnesium carbonate is obtained.

$$Mg(OH)_2 + CO_2 \quad MgCO_3 + H_2O$$

From this salt a new salt and carbonic acid could be obtained by treatment with an acid. Thus nitric acid gives the salt magnesium nitrate. The carbonic acid, produced simultaneously decomposes to carbon dioxide and water.

$$MgCO_3 + 2HNO_3 = Mg(NO_3)_2 + \underset{\displaystyle CO_2+H_2O}{\overset{\displaystyle H_2CO_3}{\downarrow}}$$

The following conversions could be effected by the above reactions.

$$Mg \rightarrow MgO \rightarrow Mg(OH)_2 \rightarrow MgCO_3 \rightarrow Mg(NO_3)_2$$

Take a non-metal, for example, sulphur. The following changes could be effected successively.

$$S \rightarrow SO_2 \rightarrow Na_2SO_3 \rightarrow H_2SO_3 \rightarrow SO_2$$

It is possible to obtain from a salt an acid (hydrochloric acid) by the action of sulphuric acid, or a base (cupric hydroxide) by the action of caustic soda.

$$\underset{\text{(Salt)}}{CuCl_2} + H_2SO_4 = CuSO_4 + \underset{\text{(Acid)}}{2HCl}$$

$$CuCl_2 + 2NaOH = 2NaCl + \underset{\text{(Base)}}{Cu(OH)_2}$$

Other metals and non-metals also show similar series of conversions from simple substances to oxides, bases, acids and salts ; from salts to acids and bases ; from acids and bases to oxides, and so on.

Metals and non-metals give oxides by their oxidation in air or oxygen. From oxides other classes of compounds may be obtained. Acids interact with bases to give salts. Acids and bases can be obtained from interaction of salts. Acids and bases give oxides by dehydration.

The relation between simple substances, oxides, bases, acids and salts, can be shown schematically as in Fig 12.

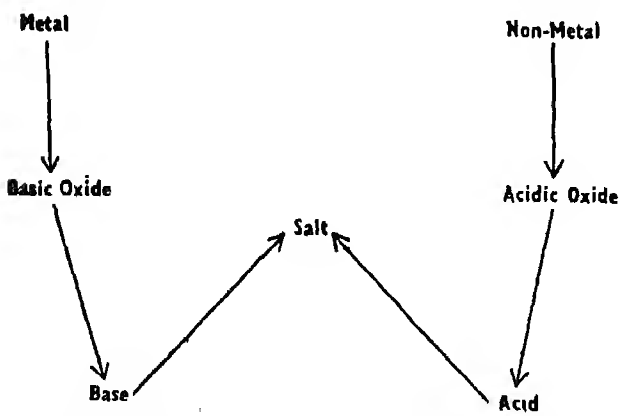


FIG. 12. Relationship between simple substances

Let us illustrate the above relationship by taking the two simple substances calcium and phosphorous (Fig. 13).

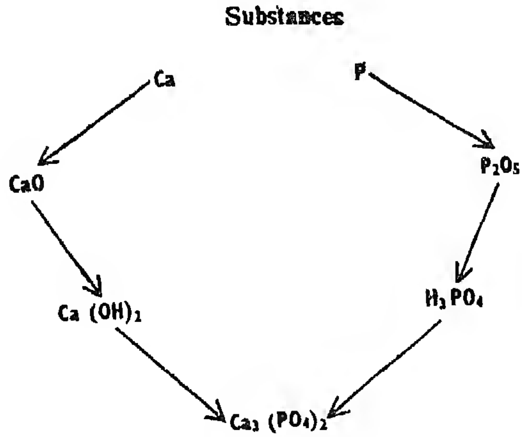


FIG. 13. Inter-relationship between simple substances and their compounds.

The many chemical conversions that take place in nature and in our practical life are based on these inter-relationships.

Practical Work No. 4

Important classes of inorganic compounds—(experimental problems).

Apparatus : Spirit lamp, test tubes with stand, retort stand with ring, asbestos gauze, porcelain dish, beakers, flasks, funnel, scissors, apparatus for preparation of gas, filter paper.

Reagents : Solutions of acids, alkalies and salts. Dry salts and solutions of litmus and phenolphthalein.

Problem 1 :

You are given two test tubes: one containing calcium oxide and the other phosphoric anhydride. Label the test tubes.

Problem 2 :

You are given three test tubes, one of which contains a solution of an acid, the other an alkali and the third a salt. Determine which substance is present in which test tube.

Problem 3 :

Prepare copper hydroxide from copper oxide and separate it from the reaction mixture.

Problem 4 :

You are given solutions of copper sulphate and caustic soda. Prepare copper oxide from these and prove that it is a basic oxide.

Problem 5 :

You are given the following substances : Na_2CO_3, $Ca(OH)_2$ and HCl. Using these prepare $CaCO_3$.

Problem 6 :

You are given the following substances : H_2SO_4, $Ba(NO_3)_2$ and Zn. Prepare a solution of zinc nitrate and separate it from the other substances.

Problem 7 :

Prepare a solution of KOH by making use of the reaction of double decomposition and the process of decantation.

Problem 8 :

You are given the following substances : $Ba(OH)_2$. $CaCO_3$ and HCl. Prepare barium carbonate and separate it from the reaction mixture.

Problem 9 :

Prepare a solution of zinc chloride using the following substances : Zn,$BaCl_2$ and $CuSO_4$.

Problem 10 :

Prepare carbonic anhydride from potassium carbonate and prove that it is an acidic oxide.

Problem 11 :

Prove experimentally the basic nature of magnesium oxide.

Problem 12 :

Prepare $Zn(OH)_2$ using Zn and NH_4OH Separate it from the reaction mixture.

TABLE

Names of the salts of some important acids

Name and formula of acid	*Valency and formula of the acid radical*	*Name of the corresponding salt*	*Formula and name of some of the salts*
Hydrochloric acid, HCl	I, $-Cl$	Chloride	$MgCl_2$ (Magnesium chloride) $NaCl$ (Sodium chloride)
Hydrogen sulphide, H_2S	II, $=S$	Sulphide	ZnS (Zinc sulphide) FeS (Ferrous sulphide)
Nitric acid, HNO_3	I, $-NO_3$	Nitrate	$Ba(NO_3)_2$ (Barium nitrate) KNO_2(Potassium nitrate)
Carbonic acid, H_2CO_3	II, $=CO_3$	Carbonate	Na_2CO_3 (Sodium carbonate) $CaCO_3$(Calcium carbonate) $NaHCO_3$(Sodium bicarbonate)
Sulphurous acid, H_2SO_3	II, $=SO_3$	Sulphite	K_2SO_3(Potassium sulphite) Na_2SO_3(Sodium sulphite) $NaHSO_3$(Sodium bisulphite)
Silicic acid, H_2SiO_3	II, $=SiO_3$	Silicate	$CaSiO_3$(Calcium silicate) Na_2SiO_3(Sodium silicate)
Sulphuric acid, H_2SO_4	II, $=SO_4$	Sulphate	$CuSO_4$(Copper sulphate) $ZnSO_4$(Zinc sulphate) K_2SO_4(Potassium sulphate) $KHSO_4$(Potassium bisulphate)
Phosphoric acid, H_3PO_4	III, $\equiv PO_4$	Phosphate	$AlPO_4$(Aluminium phosphate) Na_3PO_4(Trisodium phosphate) Na_2HPO_4(Disodium hydrogen phosphate)

Questions and Exercises

1. How can you bring about the following conversions ?

   (*a*) $P \rightarrow P_2O_5 \rightarrow H_3PO_4 \rightarrow Ba_3(PO_4)_2$

   (*b*) $S \rightarrow SO_2 \rightarrow H_2SO_3 \rightarrow BaSO_3 \rightarrow SO_2$

   (*c*) $Al \rightarrow AlCl_3 \rightarrow Al(OH)_3 \rightarrow Al_2O_3$

   Write the equations of the reactions

2. Give the equations of the reactions of preparation of sodium hydroxide by (*a*) reaction of combination and (*b*) reaction of double decomposition.

3. How many grams of nitric acid are necessary for the complete neutralization of a solution containing 14 8 grams of calcium hydroxide ? (**Ans.** 25.2 g)

4. A solution contains 28 g of caustic potash. Which acid, nitric, sulphuric or hydrochloric should be taken for its complete neutralization so that the weight of the acid taken is the least ? (**Ans.** HCl)

5. How many litres of carbon dioxide are required to get a salt, by reacting with a solution containing 16 grams of caustic soda ? (1 litre of carbon dioxide, under normal conditions weighs 1.96 g) (**Ans** 4.5 litres)

6. How can you prepare magnesium sulphate with the help of (*a*) the reaction of combination (*b*) the reaction of double decomposition,and (*c*)the reaction of displacement ?Write the equations of the reactions

7. How many grams of iron are necessary for the displacement of 12.8 grams of copper from a solution of blue vitriol ($CuSO_4$) ? (**Ans.** 11.2 g)

8. How many grams of ferrous oxide were spent on interaction with a solution of hydrochloric acid if 3.25 g of ferrous chloride is obtained in the process ? (**Ans.** 1 6 g)

9. By interaction of a solution of potassium carbonate with barium hydroxide, 3.94 g of a precipitate was obtained. What was the weight of each of the initial substances taken ? (**Ans.** 34 2 g and 27.6g)

10. Write down the equations of the reactions of the various ways of preparing aluminium chloride.
11. Starting from zinc oxide how can you obtain zinc hydroxide. Give equations for the reactions involved.
12. Write down the equations of the reactions as a result of which we can obtain barium carbonate.
13. How many kilograms of calcium oxide are needed to prepare 1000 kg of slaked lime ? (Ans. 756.7 kg)
14. Is it advisable to keep solutions of salts of (*a*) calcium (*b*) copper and (*c*) mercury, in aluminium vessels ? Explain why ?
15. How many kilograms of sodium nitrate can be obtained by the interaction of nitric acid with 20 kilograms of caustic soda ? (Ans. 42.5 kg)
16. How can you obtain iron oxide from iron nitrate ? Write the equation of the reaction.
17. How can magnesium chloride be obtained from magnesium nitrate ?

CHAPTER II

Chemistry in Agriculture

FOOD is most important in our life and agriculture is the biggest occupation of our country.

In a country like ours where the population is increasing fast the production of food should also be accelerated so that there is enough for all. Biology has contributed by producing better breeds of plants and animals, better seeds or grafts, and disease resisting but high yielding breeds of plants and animals. Chemistry plays a very important role in the development of agriculture and in the production of more and better food.

Plants are grown in soil, from which they take up the nutrients required for their growth. Thus, by the growth of plants the soil is constantly getting depleted of these nutrients. Also, all plants do not require the same kind of nutrients.

How do we know whether any given soil contains the required nutrients for any particular crop and that too in sufficient quantities? This is found out by an analysis of the soil in special soil testing laboratories.

14. Types of Soils—Acidic or Alkaline

Some soils have alkali in them and hence are either non productive or less productive. Some soils on the other hand may be acidic in character and less productive also. Good soil should be neither too much acidic nor alkaline.

Experiment : Take the given samples of soil in separate test tubes or beakers. Add some water in each and stir well. Then filter the contents and test each separately with blue and red litmus

papers. Observe and draw your conclusion about the type of the soil samples given.

Soil testing laboratories not only find the acidic or alkaline nature of soils but also make detailed analysis of all constituents.

It is known that there is as much as 12 million acres of *usar* or alkaline land in India. By suitable chemical treatment, like the application of gypsum, these alkali soils can be reclaimed for food production.

To aid gypsum in the reclamation of alkali soils, a green manure crop of "dhaincha" (Sesbania asculeata) is sometimes used.

Acid soils are treated with lime to neutralize the acidity.

The analysis of the soils helps the farmer to determine the nature of the soil and the treatment required to make up for the defects that it may have.

15. Chemistry in Areas of Agriculture

All plants require some elements in varying amounts for their growth (the total number may go up to 22 elements). Except carbon, hydrogen and oxygen all the other elements are absorbed from soil, fertilizers and manures. The most important amongst these elements are nitrogen, phosphorus and potassium which are required in large amounts by plants. These are called *primary nutrients.* The other essential elements are called secondary nutrients. The *micronutrients* are used only in very small quantities by the plants Different amounts of essential elements are needed by plants during different stages of their growth.

TABLE

Essential nutrient elements for plant growth and their sources

From air and water	*From soil, fertilizers and manures*		
	Primary nutrients	*Secondary nutrients*	*Micro-nutrients*
Carbon	Nitrogen*	Calcium	Copper
			Manganese
			Cobalt
Hydrogen	Phosphorus	Magnesium	Zinc
Oxygen	Potassium	Iron	Boron

*Leguminous plants obtain a part of their nitrogen from the nitrogen of the air : all other plants obtain nitrogen only from the soil.

The deficiency of essential elements in the soil is made good by addition of manures and fertilizers to the soil. A manure is a natural substance obtained by the decomposition of animal excreta or plant residues.

A chemical fertilizer is a salt containing certain necessary elements for the soil.

Fertilizers : Depending upon the nourishing elements, K, N and P the fertilizers are divided into the following important groups :

(*a*) potassium fertilizers;

(*b*) nitrogen fertilizers and;

(*c*) phosphorus fertilizers.

Very often they are salts of nitric, phosphoric and other acids. Each of these fertilizers contains one of the above mentioned nutrients. These are also complex fertilizers which contain two or even three of the nourishing elements. Often different fertilizers are mixed together. Such mixed fertilizers are more effective than the individual fertilizers used separately. You have already seen that such substances like lime and gypsum are employed to treat soils to make them neutral. These substances are often called *indirect* fertilizers, as they increase the effectiveness of fertilizers.

Increasing amounts of chemical fertilizers are being manufactured in our country to meet their growing demand. In the year 1965 we produced 243,000 tons of nitrogenous fertilizers and imported 104,144 tons of the same.

The main centres of fertilizer production in India today are at Sindri in Bihar, Alwaye in Kerala, Nangal in Punjab, Rourkela in Orissa and Gorakhpur in U P. The projected fertilizer factories are at Bombay in Maharashtra, Namrup in Assam, Durgapur in Bengal and Cochin in Kerala When all these factories start production in full swing, they are expected to go a long way to meet our fertilizer demands.

Control of weeds and pests: Another area of agriculture in which chemistry plays a part is in the control of weeds and pests. A weed is an unwanted plant that grows with a crop. The manures or fertilizers containing the various nourishing elements that are so essential for plant growth, are also helpful to the weeds which are useless to us. By taking away a lot of nutrients added to the soil, they increase the cost of cultivation. They also hinder the growth of useful plants. Hence, to increase the yield of crops it is very necessary to remove these weeds.

A good weed-killer should be cheap, effective, easy to apply in solution and not be harmful to man or animal. Since all these condi-

tions are satisfied by the chemical 2-4-D (2, 4 dichlorophenoxy acetic acid) it has become a popular weed-killer.

Chemistry has been used in agriculture to control pests. It is estimated that 4 rats use up as much grains as one man and since the rodent population in India is about 4 times that of the human population you can understand the tremendous loss caused by these pests. Besides, there are insect pests too like locusts, which cause extensive damage to crops.

Pests can be controlled by using chemicals in different forms —dusts, sprays or as gas—depending on the kind of pests to be controlled. Some of the common chemicals used for pest control are D.D.T., B.H.C and Zinc phosphide.These are called insecticides or pesticides.

Growth Stimulators :—The science of chemistry not only helps in providing nourishment to plants but also in activating the whole process of their development. For this purpose certain chemicals known as *stimulators of growth or growth promoting substances* are used. They do not serve as nourishment themselves but they influence the whole process of the growth and development of plants in different stages : for example, in the sprouting stage, in the flowering stage, during the development of the root system and during the development of fruits.

Ethylene is a gas that increases the development of fruits.

Note : Chemicals used for pest control are poisonous and they must be applied with extreme care. All the directions given on the label on the container, must be strictly followed. Insecticides must be kept in closed and well labelled containers in a dry place, where they may not contaminate food and where children cannot reach them.

Questions

1. Mention how chemistry is important to agriculture.
2. Why do plants need fertilizers ?

3. What are the main nourishing elements, required by plants ?
4. What are primary nutrients ? Give examples.
5. What are secondary nutrients ? Give examples.
6. What are micro-nutrients ? Give examples.
7. How are fertilizers classified ?
8. What are indirect fertilizers ?
9. What are the necessary attributes of a good weed-killer ?
10 What is an insecticide ? Give examples.

CHEMICAL FERTILIZERS

You have seen how important is the use of mineral fertilizers in modern agriculture practices. For the proper use of mineral fertilizers, one should know their composition, properties and the correct methods of using them. One should also know how to distinguish one fertilizer from the other with the help of chemical reactions.

16. Potassium Fertilizers

Potassium fertilizers are used to assist the formation of starch, sugar, proteins, fats and other substances in the plants. The shortage of potassium fertilizers in soil decreases the yield, and the plants are weakened and become diseased. The salt potassium chloride (KCl) finds more widespread use as a potassium fertilizer.

Experiment : Observe the given sample of potassium fertilizer. How does it look ? Test its solubility in water.

The potassium chloride fertilizer is a white or greyish powder, very soluble in water. It is not ordinarily *hygroscopic* (does not absorb water vapours from air) but on keeping for a long time in open air it gets sticky.

Identification of Potassium sulphate —K_2SO_4

(a) Detection of potassium

Experiment . Clean a graphite stick from a pencil. Dip it in concentrated hydrochloric acid and after that touch it with a potassium fertilizer. Introduce the tip of the graphite stick with the salt adhering to it to the top part of the flame of a spirit lamp or a burner. Repeat the same experiment with other potassium salts also.

Observe the colour of the flame with and without a blue glass.

The violet colour (characteristic colour) of the flame confirms the presence of potassium compounds. The flame appears crimson through the blue glass.

(*b*) Detection of the chloride radical

Experiment : Take a pinch of potassium chloride and dissolve it in 3 ml of distilled water in a test tube. To the transparent solution add a few drops of silver nitrate solution and observe. Now add a few drops of ammonium hydroxide and observe again

A white curdy precipitate of silver chloride settles down.

$$KCl + AgNO_3 = KNO_3 + AgCl\downarrow$$

The precipitate dissolves in a solution of ammonium hydroxide.

Potassium sulphate also belongs to the group of potassium fertilizers Sometimes it is used in place of potassium chloride because the chloride radical is harmful to some plants

Identification of Potassium sulphate —$K_2\ SO_4$

(a) Detection of potassium

It can be detected as in the previous case by the flame test as the colour of flame for all salts containing potassium is *violet.*

(*b*) Detection of the sulphate radical

Experiment : Take a small quantity of potassium sulphate and dissolve it in 3-5 ml of water. Add a few drops of barium chloride

solution and observe Add to the precipitate 1-2 ml of concentrated nitric acid and observe again.

A white precipitate of barium sulphate is obtained.

$$K_2SO_4 + BaCl_2 = 2KCl + BaSO_4\downarrow$$

The precipitate does not dissolve in concentrated nitric acid.

Wood ash contains 15 to 18 percent of potash (Potassium carbonate (K_2CO_3) and other potassium compounds. Hence, it is an important potassium fertilizer.

Task

Prepare a plan for identifying potassium carbonate and perform the required experiments.

Questions

1. Calculate the percentage of potassium in pure
   (*a*) KNO_3 (*b*) KCl (Ans (*a*) 38.7% (*b*) 52.2%)

2. A 0.4 percent solution of KCl is used in potato cultivation. How many grams of potassium chloride are necessary for the preparation of 10 kilograms of such a solution ?
   (Ans. 40 grams)

3. The following mineral fertilizers are given to you in test tubes :
   (*a*) KCl, (*b*) K_2SO_4.
   How will you distinguish one from the other ?

4. If 5 kg of combined potassium is removed from the soil by crops, how much of (*a*) KCl, or (*b*) KNO_3 should be added to the soil to replenish its potassium content ?
   (Ans. (*a*) 9.5 kg, (*b*) 12.9 kg)

5. The ash of the stems of sunflower plants contains 28% of potash (Potassium carbonate, K_2CO_3). How much ash by weight should be put in the soil in place of 1 ton of potassium chloride to provide an equivalent amount of potassium ?
   (Ans. 3.3 tons)

17. Nitrogen Fertilizers

Nitrogen is needed for the formation of proteins. Without proteins life would be impossible for both animals and plants. Some nitrates and a few other nitrogenous compounds like ammonium sulphate and urea are used as nitrogen fertilizers.

Sodium nitrate (Chilè Saltpetre)-$NaNO_3$

Experiment : Take the given sample of sodium nitrate in a test tube. Observe its state, colour and appearance. Test if it dissolves in water.

In the pure form, *Chile Saltpetre* is a white crystalline substance very soluble in water. It is prepared by the interaction of sodium carbonate and nitric acid.

$$Na_2CO_3 + 2HNO_3 = 2NaNO_3 + H_2O + CO_2$$

Commercial Chile saltpetre usually contains different impurities which impart a grey or a greyish-yellow colour to it. It contains 16.1% of nitrogen and is used as a nitrogen fertilizer.

Identification of sodium nitrate (*Chile saltpetre*)—$NaNO_3$

Detection of Sodium

Experiment : Take a platinum wire mounted on a glass rod. To clean it, dip it in concentrated hydrochloric acid and dry it on a flame. Dip the wire again in acid and then touch it to the dry sodium nitrate. Hold the wire (with some crystals of the salt adhering to it) in the top part of the flame of a spirit lamp or burner. What do you observe ?

(This experiment may be performed with a fresh clean graphite stick from a pencil, in place of the platinum wire).

Repeat the experiment with another sodium salt, say, sodium chloride. What do you see in this case ?

Experiments with different salts of sodium show that in all the cases the flame is coloured golden yellow and that this is a characteristic property exhibited by all compounds of sodium.

Detection of the acid radical (nitrate)

Take a piece of charcoal and make one end of it red-hot. Sprinkle some sodium nitrate powder over it ?

What do you observe ?

Repeat this with calcium nitrate and observe.

The property of producing sparkling flashes with a piece of red-hot charcoal is shown by all nitrates.

Calcium nitrate—$Ca(NO_3)_2$

Calcium nitrate is another compound used as a nitrogenous fertilizer.

It is produced as granules or in the form of white scales which are soluble in water. It contains 15-16 percent of nitrogen, and is *hygroscopic* (i.e. absorbs water on standing in air and gets sticky). Hence it is necessary to keep calcium nitrate in special containers and in dry conditions.

Calcium compounds can be detected easily as they impart a brick-red colour to the flame.

Ammonium nitrate—NH_4NO_3

It is one of the main nitrogen fertilizers. Observe a sample of ammonium nitrate and test its solubility in water It is a solid crystalline substance, white or yellowish white in colour and is very soluble in water. Ammonium nitrate is very rich in nitrogen (34.35%). It is a salt of nitric acid.

In the composition of ammonium nitrate there is a group of atoms (NH_4) instead of a metal. This group is called *ammonium group or ammonium.* This group ammonium exists only in the composition of a compound as a monovalent radical. Just as there are the metals Na and K in the bases NaOH and KOH, there is the group NH_4 in the base NH_4OH. This base is called ammonium hydroxide Ammonium hydroxide decomposes very easily with the evolution of ammonia gas which is easily recognised by its characteristic smell.

$$\underset{\text{Ammonium hydroxide}}{NH_4OH} = \underset{\text{Ammonia}}{NH_3\uparrow} + \underset{\text{Water}}{H_2O}$$

Does it behave like other bases ?

Experiment : Take 2-3 ml ammonium hydroxide solution in a test tube. Smell it carefully. Test a drop with litmus paper to find out its nature whether acidic or basic.

Add to this solution drop by drop, dilute nitric acid till it is neutralized. Evaporate carefully a few drops of the clear solution to see the deposit of salt obtained.

Like all other bases, ammonium hydroxide interacts with acids to form salts.

$$NH_4OH + HNO_3 = H_2O + \underset{\substack{\text{Ammonium nitrate} \\ \text{(White salt)}}}{NH_4NO_3}$$

Ammonium hydroxide also gives salts with other acids. These salts are called ammonium salts

On standing in air, ammonium nitrate in the form of fine crystals absorbs water from air and becomes sticky. Hence it is necessary to keep it in dry conditions. Ammonium nitrate is produced industrially as a powder or in the form of granules.

Identification of ammonium nitrate—NH_4NO_3

Experiment : Take a pinch of ammonium nitrate in a test tube and add 2-3 ml of a 20 percent solution of caustic soda. Heat for a little while. Smell carefully the gas evolved. Hold a piece of moist red litmus paper over the mouth of the test tube. Note the change in the colour of the paper. Why the paper changed its colour to blue ?

$$NH_4NO_3 + NaOH = NaNO_3 + NH_4OH$$
$$NH_4OH = NH_3\uparrow + H_2O$$

Whenever an ammonium salt is heated with alkali, ammonia gas is evolved.

Test the acid radical (nitrate) in ammonium nitrate by the method you have already learnt and record your results.

Ammonium sulphate— $(NH_4)_2SO_4$

Another widely used nitrogen fertilizer is ammonium sulphate. Observe the given sample of ammonium sulphate and test its solubi-

lity in water. It is a fine crystalline powder of white or grey colour. It is very soluble in water and does not become sticky on standing. Calculate the percentage of nitrogen in ammonium sulphate (21%).

Identification of ammonium sulphate – $(NH_4)_2SO_4$

Experiment: Detect the ammonium group in the sample of ammonium sulphate by the method already learnt.

For detecting the acid radical, dissolve about 0.5 gram of ammonium sulphate in 3-4 ml of water in a test tube. Add to this solution 5-6 drops of a solution of barium chloride. What do you observe ? Add to the precipitate 1-2 ml of concentrated nitric acid. Does the precipitate dissolve ?

$$(NH_4)_2SO_4 + BaCl_2 \rightarrow 2NH_4Cl + \underset{\text{(insoluble)}}{BaSO_4\downarrow}$$

Observe what happens when you repeat the experiment with potassium sulphate or sodium sulphate in place of ammonium sulphate. An insoluble precipitate of $BaSO_4$ is formed.

Another important nitrogen fertilizer is *urea* Like other fertilizers urea is also manufactured by artificial means. The composition of urea is $CO(NH_2)_2$. Among all the nitrogen fertilizers, it has the highest percentage of nitrogen (46.66%).

Questions

1. which salts of nitric acid are used as mineral fertilizers? Write their formulae.
2. Calculate the percentage of nitrogen in pure ammonium nitrate. (**Ans.** 35%)
3. A 20 hectares field of wheat needs 25 kilograms of nitrogen per hectare. What quantity of ammonium sulphate would provide this amount of nitrogen ? (**Ans.** 2,327 kg)
4. What quantity of ammonium nitrate is required to be put in a field of 25 hectares of wheat, if, for each hectare, 45 kilograms of nitrogen are needed. What quantity of ammonium sulphate will be required if it is used in place of ammonium nitrate ? (**Ans.** (1) 3,214 kg of NH_4NO_3
   (2) 5,303 ,, ,, $(NH_4)_2SO_4$)

5. What are the similarities and differences between the composition of ammonium sulphate and ammonium nitrate? How can you distinguish one from the other?

6. Write the equations of the interaction of ammonium nitrate (NH_4NO_3) and ammonium sulphate, $(NH_4)_2SO_4$, with an alkali.

7. For an average yield of wheat from a field of one hectare for one season, 75 kilograms of nitrogen are required. How many kilograms of ammonium nitrate should be put into a field of 1000 hectares to get a similar yield of wheat if only 20% of nitrogen is already available from the natural processes (decaying of plants, growing of micro-organisms, etc.)? (Ans. 171.4 tons)

18. Phosphatic Fertilizers

Phosphorus is one of the main elements needed by a plant. It increases the yield of food grains. It increases the sugar content in the beet-root and the starch content in potatoes and corn. It also improves the quality of the fibres of flax and accelerates the ripening of plants A decrease of phosphorus content in the soil reduces the absorption of nitrogen and other elements by plants

The basic raw material for production of phosphatic fertilizers is the natural mineral *phosphorite.* It is a salt of phosphoric acid and its composition is $Ca_3(PO_4)_2$.

Almost all of the phosphatic fertilizers used in our country are superphosphates. Sometimes rock phosphate, bone meal and basic slag are also used. *Phosphorite,* obtained from the natural mineral, is a very fine powder of dark grey colour It is insoluble in water and therefore the phosphorus contained in it is not readily available to plants. Hence the powder is used in some acidic soils, which convert the phosphoric powder into a soluble salt of phosphoric acid

The effect of phosphorite is increased when mixed with other fertilizers, e.g., ammonium sulphate, manure and compost.

Superphosphate [$Ca(H_2PO_4)_2 + 2CaSO_4$]

The superphosphates are also obtained from the mineral phosphorite. The cleaned and powdered mineral is treated with dilute sulphuric acid. The following reaction takes place:

$$Ca_3(PO_4)_2 + 2H_2SO_4 = Ca(H_2PO_4)_2 + 2CaSO_4$$

The mixture of $Ca(H_2PO_4)_2$ and $2CaSO_4$ is called *superphosphate.* Usually superphosphate contains some water.

It is a white or grey powder and contains about 9% of phosphorus The essential part of superphosphate is the *acid calcium phosphate*—$Ca(H_2PO_4)_2$. Such acidic salts are more soluble in water than the corresponding normal salts. Thus, these are easily absorbed by plants. The other component, calcium sulphate, is almost insoluble in water. Like many other fertilizers, superphosphate is often used in granular form. Granulated fertilizers are easily taken up by the plants.

Identification of superphosphate -($Ca(H_2PO_4)_2 + 2CaSO_4$)

To identify a sample of superphosphate one has to detect the presence of acid calcium phosphate in the sample as it is the essential part of the fertilizer.

Experiment : Shake a little of the given sample with a few ml of water in a test tube. Allow it to stand for some time, and then filter. Add 1-2 ml of a solution of silver nitrate to the transparent filtrate and observe.

The appearance of a yellow precipitate shows the presence of acid calcium phosphate The calcium sulphate remained as a residue on the filter paper while the filtrate contained some acid calcium phosphate, the latter being slightly soluble in water.

Questions

1. What are phosphatic fertilizers ? Why are they required by plants ?
2. What is phosphorite (powder) ? What are its properties ?
3. What is superphosphate fertilizer ? How is it produced ?

4. How can superphosphate be distinguished from potassium chloride fertilizer ?
5. How much superphosphate can be obtained from 10 tons of natural phosphorite which contains 20 percent of impurities ? (Ans. 13 tons)

19. Micro-Fertilizers (Trace Elements)

You know that for our growth and health we need some elements like iron, copper, calcium etc. in traces, which we get through various food materials, tonics and vitamins. Similarly for their growth, the plants need some elements in traces. Fertilizers containing such elements are called *micro-fertilizers* or trace elements. Micro-fertilizers not only increase the yield of crops but also increase the resistance of plants against different diseases.

Boron increases the quantity of sugar and vitamins in fruits and vegetables. The shortage of boron fertilizers in soil causes the delay in the growth of the plant and the decay of the roots, e.g. in sugar beet.

Due to shortage of copper in soil, the plants get the disease of *white plague*. The edges of leaves develop a white colour and get dry.

Manganese increases sugar in beet-roots, starch in potatoes and corn and improves the quality of wheat seeds.

The main micro-fertilizers are

1.	Boric acid	H_3BO_3—white crystals.
2.	Borax	$Na_2B_4O_7.10H_2O$—white crystals
3.	Blue vitriol	$CuSO_4.5H_2O$—blue crystals
4.	Manganese sulphate	$MnSO_4.4H_2O$—rose crystals.

All of these fertilizers are very soluble in water.

All of them are put in the soil in the form of solutions, containing 0·2—0·6 gram of fertilizer per litre of water. They are also applied in small quantities mixed with superphosphate, potassium nitrate and other fertilizers.

The aqueous solutions of micro-fertilizers are used for spraying of plants and for moistening seeds before sowing.

20. Mixed Fertilizers

Fertilizers having only one nourishing element in their composition are called *single* fertilizers. Often not less than two nourishing elements are put into the soil and for this purpose mixed fertilizers are used. They contain 2 or 3 nourishing elements and sometimes trace elements. Depending on the presence of the important elements, such fertilizers are called NK,NP,PK and NPK.Fertilizers of the NPK—type are called *complete* fertilizers. Production, transport and usage of mixed fertilizers are cheaper than an individual fertilizer used separately. The former are more effective also.

KNO_3 belongs to the fertilizers of the NK - type. It contains 2 nourishing elements—nitrogen and potassium. Cheap mixed fertilizers of NP - type are prepared by mixing simple superphosphate with ammonium sulphate.

Among the mixed fertilizers, those of the NP - type are the more important ones. Ammonium salts of phosphoric acid belong to this type of fertilizers, like *Ammophos* (Ammonium hydrogen phosphate). It has the formula $(NH_4)H_2PO_4$. Ammophos is obtained either as a white powder or in the form of granules. It is very soluble in water and is non-hygroscopic.

Animal bones are also used as a phosphatic fertilizer in the powdered form. They contain mainly calcium phosphate—$Ca_3(PO_4)_2$.

One of the important mixed fertilizers is *nitrophosk*. This is a complete fertilizer containing nitrogen, phosphorus, and potassium (NPK). 'Nitrophosk' is prepared by melting together ammophos, potassium sulphate and ammonium nitrate

Different composts prepared by mixing mineral fertilizers with the so-called *local fertilizers* belong to the category of mixed fertilizers Examples of local fertilizers are dung, manure and ash. Manure is a full (complete) fertilizer. But the nourishing elements present in it are insufficient. During storage, a manure loses a good amount of its nourishing elements, especially nitrogen. For better results manure is mixed with 'phosphorite flower', with superphosphate and other fertilizers.

Mixture of mineral fertilizers in the form of aqueous solutions are used for the growing of plants without soil.

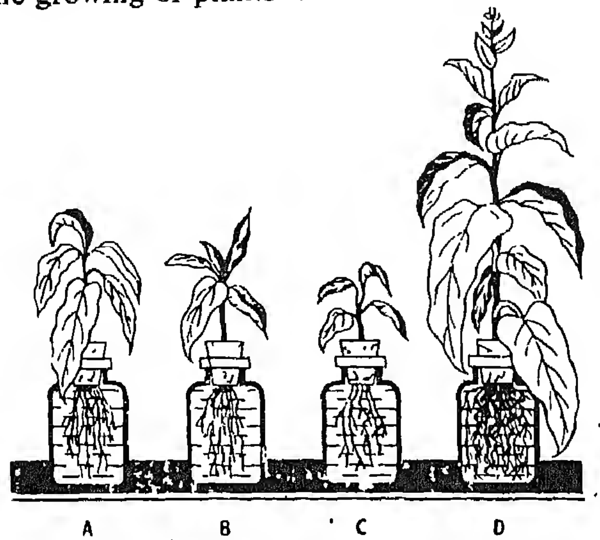


FIG. 14.

Plants can grow and develop by taking all the important nourishing elements from the specially prepared solutions or from sand, into which such a solution is poured. Such a method of growing of plants is called *hydroponics* (Fig 14) In the green houses (hot house in cold countries) and also in open air, tomatoes, cucumbers, radishes, onions and other vegetables are grown. Plants grow and develop quicker in nourishing solutions than in ordinary soil. For example, tomatoes in such conditions can give up to 6 times the normal yield in a year.

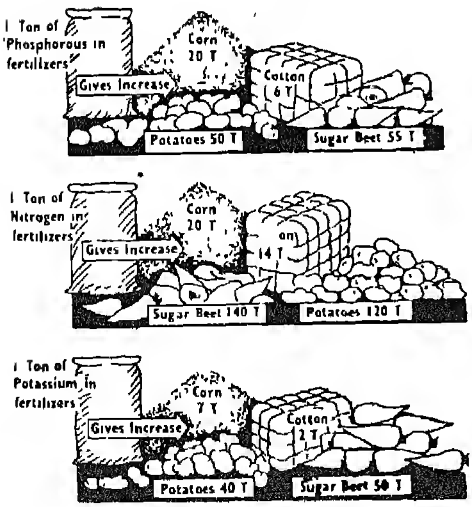


FIG. 15. Yield from equivalent amounts of different fertilizers

Yield from equivalent amount of various fertilizers is shown in Fig. 15.

Questions

1. What are trace elements and of what use are they to the plants ?
2. What are single fertilizers ? Give two examples.
3. What are complete fertilizers ? Give two examples.
4. What are mixed fertilizers ? Give two examples.
5. How can you identify potassium nitrate ?

21. Protection of Plants from Parasites and Diseases—Weed Control

For obtaining good yields not only the treating of soil and providing nourishment for plants are essential but preventing of disease of plants and avoiding weeds which remove nourishment and lower the growth of plants are also equally important Due to weeds and plant diseases, agricultural products worth crores of rupees are lost every year. Using fertilizers without destroying weeds is irrational since not only the crops but also the weeds grow better in fertilized soil.

Some of the chemicals used against weeds and parasites have been known for a long time. Recently, several newer and much more effective remedies have been discovered for the fight against weeds and plant diseases. All of them have complex compositions and difficult names and hence we shall consider only a few of them here.

Insecticides

The word *insecticide* means insect killer. Some of them kill the parasites by simple contact while others are absorbed by the plants and the plant juices are made poisonous to the insects. Insecticides are used in the form of solutions or suspensions, and sometimes as a powder. One of the widely used insecticides is *Paris green.* It is a

complex compound containing arsenic and copper. Paris green kills the parasites of trees and vegetables. It is used in the form of a suspension of 2 grams per litre of water. For fighting cereal parasites sodium fluosilicate (Na_2SiF_6) is employed. It is used either as a powder or as a solution for the treatment of young plants.

Sometimes a solution (2-3%) of barium chloride is used for fighting parasites of sugar beets and vegetables. It is a white crystalline substance very soluble in water.

One of the strongest insecticides is benzene hexachloride (B H C), $C_6H_6Cl_6$, also known as Gammaxene. Benzene hexachloride is used as a powder mixed with other substances. It kills practically all the parasites.

Remedies against plant diseases

Such remedies are called *fungicides*. Formalin in the form of a dilute solution is an example of a fungicide. Seeds are treated with this solution before they are sown. Sulphur dioxide gas is used to fumigate special store houses and green houses for plants.

Blue vitriol and lime are also used as fungicides. *Bordeaux mixture*, a blue green suspension which is prepared by mixing equal volumes of 1% solution each of blue vitriol and milk of lime is used for spraying green plants, e.g., apples and grapes etc.

Pesticides : Remedies against rodents

These are called *rodenticides.* Zinc phosphide (Zn_3P_2) is widely used for this purpose. It is a grey powder and is used, mixed with food articles, as a rodent poison.

It should be remembered that almost all the remedies are also poisonous for animals and human beings and hence they should be used with great care.

Weed-killers

Chemical preparations for removing weeds are called *weed-killers*. Many such substances are known.

One of the simplest and most widely used weed-killers is kerosene. It is used for the chemical removal of weeds from vegetable carrots, dhania (coriander) and sonf (fennel).

For killing weeds in the onion fields, calcium cyanamide ($CaCN_2$) is employed. It is a grey powder. 2-3 kilograms of this powder is required for each hectare. The leaves of the weeds fall off after treatment with this weed-killer.

Questions

1. What is an insecticide ? Name two insecticides known to you and state the methods of their use.

2. What are fungicides ? How are they used ?

3. Name a rodenticide known to you.

4. What is the role of weed-killers in agriculture ? Name two weed-killers.

5. Calculate the percentage of chlorine in benzene hexachloride. (Ans. 73.2%)

6. How much barium chloride is required for preparing 200 kilograms of a 3% solution for spraying on a sugar beet field ? (Ans. 6 kg)

7. Calculate the percentage of nitrogen in calcium cyanamide. (Ans. 35%)

8. How is Bordeaux mixture prepared ?

22. Preservation of Food

It is very well known that only fresh food is good. During storage, sometimes food loses its nourishing properties. By the action of oxygen of air the proteins and carbohydrates of food get fermented and such food becomes less nourishing. Chemists have discovered methods to preserve food, using special chemical preparations.

Precautions in the Use of Fertilizers

When you are working with mineral fertilizers and pesticides, etc., it is necessary to take some precautions during their addition to the soil.

Superphosphate is prepared in chemical plants by the action of sulphuric acid on natural phosphorite. Hence superphosphate always contains some sulphuric acid. If superphosphate dust falls on your skin, it gets red and inflamed. It is necessary to prevent eyes and breathing organs from coming into contact with superphosphate dust. During loading and other operations with superphosphate it is necessary to protect the eyes and cover the mouth and nose with special cloth.

Nitrogen fertilizers, especially salts of nitric acid, should be stored in such a way that there may be no possibility of their being contaminated with organic substances like coal, saw dust etc. Carelessness in the observance of this rule may cause fires from occasional sparks.

Ammonium fertilizers should never be stored together with lime nor should they be put in empty lime containers.

All the fertilizers and pesticides, weed-killers, etc., should be stored in specially labelled containers or packets. If the label is lost it is necessary to identify the fertilizer, and label it. When working with fertilizers, clothes covering the entire body should be used. In certain cases a gas mask should be employed. Fields treated with the pesticides etc , should be prevented from being visited by domestic animals and birds.

Practical Work No. 5

Identification of mineral fertilizers

In this work you have to identify the following fertilizers : sodium nitrate, ammonium nitrate, potassium nitrate, ammonium sulphate, superphosphate and potassium chloride.

Equipment : Spirit lamp and test tubes with holder.

Reagents : Solutions of hydrochloric and nitric acids, solution of alkali, barium chloride and silver nitrate, a piece of charcoal, freshly cleaned graphite stick from a pencil (or a platinum wire, if available).

Procedure : You are given several fertilizer packets. Each packet has a number on it, but no name. Using the methods already known to you, identify the fertilizers and tabulate your results as shown on page 79.

Note : Read the procedure for identification of fertilizers and then perform the experiments in the following order :—

1. Determine the solubility of fertilizer in water. Poor solubility and formation of a turbid solution indicates superphosphate, Confirm by carrying out the experiment with silver nitrate.

2. If the fertilizer is soluble in water, test it by red hot charcoal. If a flash appears, then it indicates that this fertilizer is a salt of nitric acid. For finding out which particular salt it is, test it in the flame and carry out the experiment with alkali. A yellow colour of the flame indicates sodium nitrate, violet colour shows potassium nitrate; evolution of ammonia on heating with alkali shows ammonium nitrate.

3. If the fertilizer is soluble in water, but does not give a flash with red hot charcoal it could be ammonium sulphate or potassium chloride. Test the solution by silver nitrate and barium chloride and carry out the experiment for the determination of ammonium group.

IDENTIFICATION OF FERTILIZERS

No. of packet	*Characteristic property by which it was identified*	*Name of fertilizer*
1		
2		
3		

Remark :

Data on properties and methods of identification of some fertilizers are given in the table on pp. 79 and 80.

PROPERTIES AND METHODS OF IDENTIFICATION OF SOME FERTILIZERS

Name of fertilizer	*Appearance*	*Solubility in water*	*Interaction with red hot charcoal*	*Interaction of water solution of the fertilizers*			*Other properties*
				with solution of alkali and heat	*with solution of barium chloride*	*with solution of silver nitrate*	
(1)	(2)	(3)	(4)	(5)	(6)	(7)	(8)
Sodium nitrate $NaNO_3$	White or greyish crystals	High	Gives a flash	Does not interact	No precipitate	Does not interact	Colours flame yellow
Ammonium nitrate NH_4NO_3	Granules or crystals of white or yellowish colour	Very high	Gives a flash and melts giving a white smoke	Evolves ammonia	"	"	xx
Potassium nitrate KNO_3	White crystals	High	Gives a flash	Does not interact	"	"	Colours flame violet
Ammonium sulphate $(NH\ )_2SO_4$	White or grey crystals	High	Darkens	Evolves ammonia	Gives white precipitate insoluble in hydrochloric acid	turns turbid	xx

(1)	(2)	(3)	(4)	(5)	(6)	(7)	(8)
Simple superphosphate	Powder or granules of grey or dark grey colour	Soluble partially	Does not melt or burn but gives out smell of burnt rubber	With excess of alkali gives a precipitate soluble in water	Does not interact but gives a slightly turbid solution	Solution becomes yellow, precipitate soluble in nitric acid	Colours flame dull red
Potassium chloride KCl	Small white or grey crystals	High	Cracks without melting and burning	Does not interact	Does not interact but gives slightly turbid solution	Gives curdy white precipitate, insoluble in nitric acid, but soluble in ammonium hydroxide	Colours flame violet

CHAPTER III

Carbon and its Compounds

Symbol—C Atomic weight—12

23. Carbon in Nature

You have learnt about certain important classes of inorganic compounds and mineral fertilizers. We shall now consider the element carbon that plays an exceedingly important role in our daily life Though it constitutes only about 0.03% of the earth's crust, still it is one of the most important chemical elements. The bodies of all organisms, plants and animals contain this element. Almost all our food products and fuels contain carbon. Carbon is present in coal and petroleum occurring in nature and in minerals like chalk and limestone. As you will study later in this chapter, even brilliant diamond is also composed of carbon atoms.

Demonstration : Take four dry test tubes, each containing a small amount of wood chips or saw dust, sugar, cotton, and wool. Keep a delivery tube fitted to a cork ready for testing any evolved gas (Fig. 16).

Warm carefully each test tube by turn and observe the black residue left in each case or any other change. Then heat more strongly each test tube separately and let the evolved gas pass into a test tube containing lime water. Observe the change in lime water and also the residue left in each test tube.

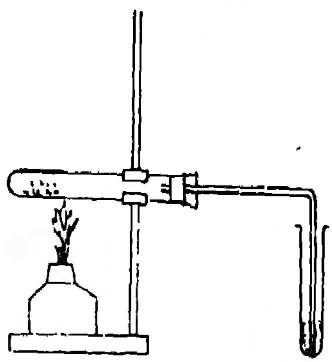


FIG. 16.
Heating of carbonaceous matter

In all these cases the black residue consists of carbon and in each case carbon dioxide is obtained on stronger heating in air, together with certain other substances like water. The residue left as ash contains other elements.

24. Modification of Carbon—Allotropy

Graphite

You must have seen the greyish-black rod inside a pencil. It is made of graphite (from the Greek word 'grapho' meaning to write).

Graphite occurs in nature. It is a soft dark grey crystalline substance having a soapy touch. It is very stable to heat, melting only at a temperature of over 3700°C. It is also a very good conductor of electricity. These properties enable graphite to find a wide use in industry. Thus, crucibles for melting metals, vessels for conducting chemical reactions at high temperatures, electrodes and other articles are made from graphite. It is also used as lubricant for machine parts.

Some types of pencils (e.g. drawing) are made of almost pure graphite. For making hard pencils other substances are added to graphite.

Graphite burns in pure oxygen. The *only gas that* it gives is *carbon dioxide*. Now, carbon dioxide is obtained by burning of any substance containing the chemical element carbon. Therefore, graphite is also a simple substance comprising of the chemical element carbon.

Diamond

Another naturally occurring form of carbon is diamond. Diamond does not look like graphite. Pure diamond is crystalline. It is colourless, transparent and has its own brilliance.

Diamond is an expensive stone because it is found only at a few places in nature. When cut and polished it is used in jewellery due to its brilliance.

The discovery of diamond in India and its cutting was made several thousand years ago But now Indian production is negligible and is limited to Panna mines in M.P. and Golkunda mines in Mysore. Some diamonds produced in India have been historically famous like Koh-i- noor, Moghul, Darya-i-noor, Nizam etc.

Diamond is the hardest among all the known natural substances. This property is used in making borers and cutters needed for the production of metallic parts of machines and for drilling holes in the hard crust of the earth (See Fig. 17, a diamond drill). Small diamonds are also used for cutting glass.

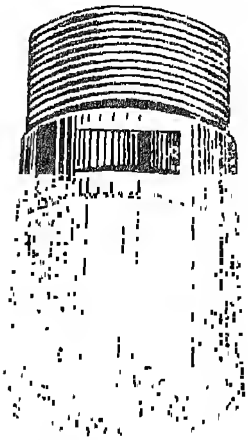


FIG. 17. A diamond drill

On heating to redness, diamond burns in oxygen *giving only carbon dioxide.* Furthermore, by heating to red hot in the absence of air, diamond has been converted into graphite. So diamond also is a simple substance consisting of the chemical element carbon. Thus, atoms of carbon could form several simple substances, which under certain conditions are inter-convertible.

The ability of elements to exist in a free state in the form of two or more simple substances is called allotropy.

Simple substances which are obtained from one and the same chemical element are called allotropic modifications.

Graphite and diamond are allotropic modifications of the chemical element carbon. The different properties of diamond and graphite

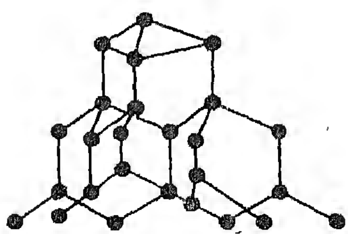


FIG. 18 (a) Distribution of carbon atoms in space in diamond

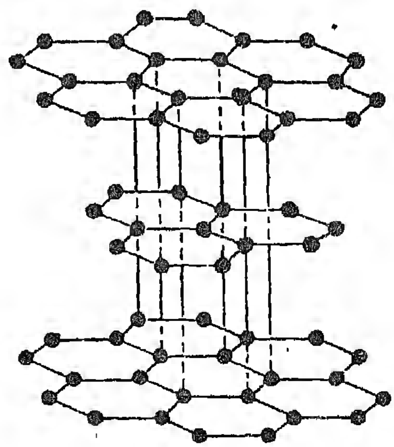


FIG. 18 (b) Distribution of carbon atoms in space in graphite

can be explained by the difference in the modes of distribution in space of the carbon atoms in the two substances. (See Figs. 18 (a) and (b))

Besides carbon many other elements also exhibit allotropy.

Amorphous Carbon

The ordinary charcoal is really a mixture whose major constituent is carbon. It is not considered an allotropic modification of carbon, for the position of carbon atoms in it is similar to graphite. It differs from graphite in having a porous and amorphous structure. Coke, bone-charcoal and soot are other forms of amorphous carbon.

Questions

1. What is the importance of carbon compounds in nature ?

2. How can you prove that diamond and graphite are simple substances consisting of the chemical element carbon ?

3. What is allotropy ? Give example.

4. How can you explain the different properties of diamond, graphite and charcoal ?

25. Preparation and Properties of Charcoal

Charcoal is obtained by burning wood in a limited supply of air. This method is, however, wasteful, since during the burning of wood a lot of charcoal is burnt and some other useful substances are also lost. To avoid such losses, wood can be decomposed thermally in a closed vessel in the absence of air. This process is called *dry distillation.*

Demonstration : Assemble the apparatus as shown in Fig. 19. Fill a third of the wide test tube with dry splinters or saw dust and fix it on a stand. Heat the test tube for some time by a spirit lamp and observe the changes. Burn the gas evolved at the mouth of the delivery tube. When the evolution of the gas ceases, put a porcelain dish under the test tube and carefully open the latter.

Pour out the liquid obtained and note that it has a distinct smell and is acidic to litmus paper. Transfer the charcoal obtained on a sheet of paper and observe it..

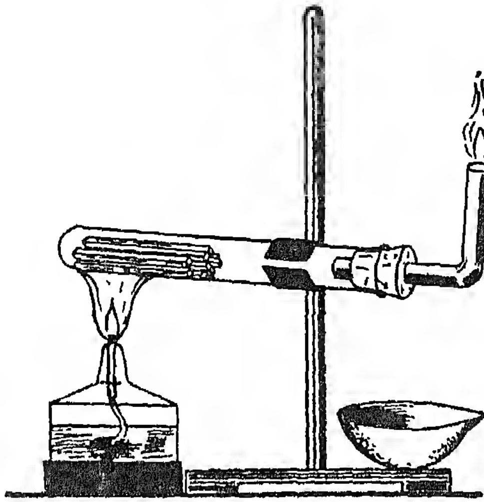


FIG. 19 Destructive distillation of wood

From the above experiment it may be observed that by the thermal decompostion of wood the following substances are obtained :

1. Combustible gas called wood gas.
2. Some acidic liquid substances; and
3 Charcoal.

The liquid obtained consists mostly of water and some useful products like acetic acid (vinegar), tar (the black floating drops), methyl alcohol (wood spirit) and several others. In industry, saw dust, splinters, logs of wood and waste timber are used for the thermal decomposition of wood. The liquid and gaseous products which are obtained in the process are used as raw materials in industry.

Charcoal obtained from wood retains the structure of wood. A piece of charcoal has numerous pores which were the vascular vessels of the wood (Fig. 20). While finely powdered charcoal sinks in water (sp. gr. 1.8 to 2.1), lumps or pieces of charcoal float on it because of its porous structure. When a piece of charcoal is kept immersed in boiling water with the help of a wire, the air escapes from the pores and is replaced by water. The piece then sinks in water.

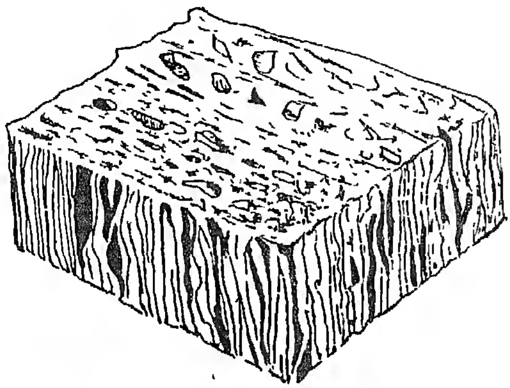


FIG 20. A piece of wood-charcoal

Demonstration : Take a stopperred bottle or flask containing reddish brown bromine vapours and put several pieces of charcoal in it. Keep the bottle stopperred for some time till the reddish brown colour disappears.

Heat the flask gently to observe the vapours of bromine being liberated from the charcoal.

Experiment : Take a small quantity of a solution of the red dye fuchsine in water and add a few charcoal pieces to the coloured liquid in the test tube. Close it with a stopper and shake several times. What do you observe now ? Has the colour changed ? If so, how ?

The property of adsorbing gases and other substances from solution is exhibited by charcoal and some other substances

The phenomenon of adsorbing gaseous or dissolved substances by the surface of a solid substance is known as adsorption. Substances on the surface of which adsorption takes place are known as adsorbents.

The ability of a solid substance to adsorb gases and dissolved substances depends on the extent of its surface area.

The porous structure of charcoal gives a large surface area. The total surface area due to the pores (capillaries) of a one gram piece of charcoal has been estimated to be between 200—1000 sq metres depending upon the type of charcoal. This big surface area of charcoal confers the important property of adsorbing gases and other substances and is used for the cleaning of solutions (de-colourization of coloured products is an example).

Charcoal is an active adsorbent. The adsorbing capacity of charcoal is increased by exposure to steam which removes the remnants of resin and tar from the surface of charcoal. Such charcoal is known as *activated charcoal*. It is used for the purification of water (Fig. 21), purification of sugar-cane juice and for the absorption of diffe-

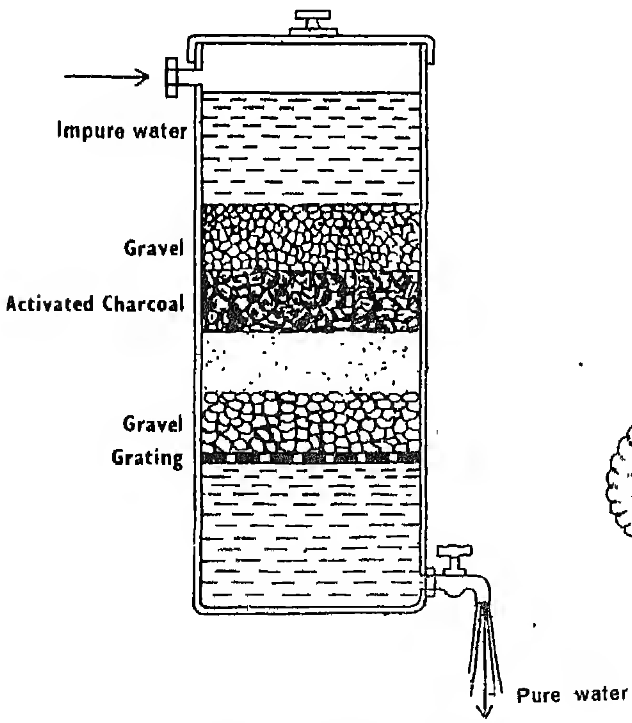


FIG. 21. Purification of water by activated charcoal filter

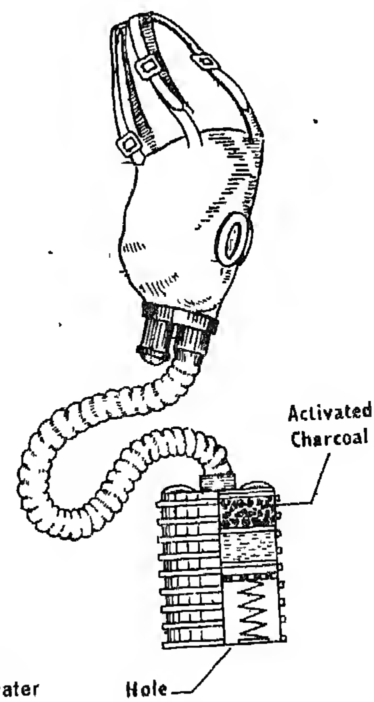


FIG. 22. A gas mask

rent gases. It is also used in gas masks for the protection of respiratory organs from poisonous gases (see Fig. 22).

Questions

1. What is meant by the dry distillation of wood ? What are the products obtained from it ?
2. Why does a piece of charcoal float on water ?
3. What is adsorption ? Give examples and state its practical applications.
4. What is activated charcoal ? What are its uses ?

26. Chemical Properties of Carbon

One of the important properties of carbon is its ability to combine with oxygen during burning. Graphite and diamond would burn only in pure oxygen but coal would burn in air. Carbon does not react with oxygen or air at room temperature. Heat is necessary to start the burning. Generally, charcoal and coke are employed for chemical reactions involving the use of carbon. On burning coal a lot of heat is evolved. Large amounts of coal are used in industry as a source of fuel. Let us see if carbon could interact also with other compounds containing oxygen, for example oxides of metals.

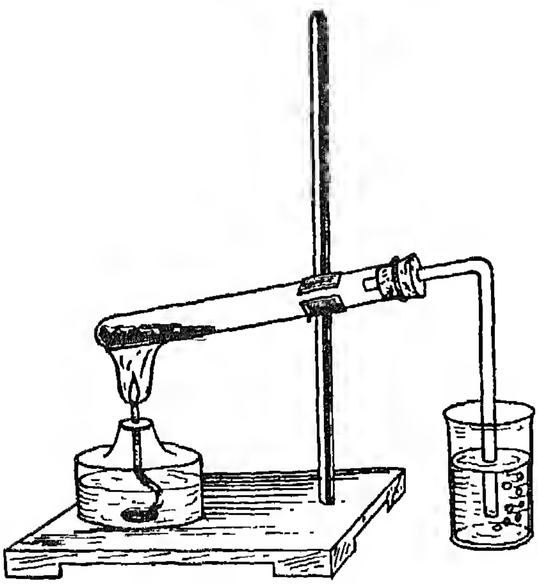


FIG. 23 Reduction of copper oxide

Demonstration : Mix some black powder of copper oxide with finely powdered charcoal. Put the mixture in a test tube fitted with a delivery tube, the end of which is dipped in a beaker containing lime water (see Fig. 23). Heat the test tube carefully in the beginning and more strongly later.

After sometime a red powder of copper appears in the test tube and the lime water turns milky.

The following reaction takes place in the test tube :

$$2CuO + C = 2Cu + CO_2$$

In this reaction copper oxide is reduced to copper; carbon being the reducing agent. Many other metallic oxides can be reduced to the corresponding metals by heating them with carbon.

Experiment : Mix a pinch of yellow powder of lead oxide with charcoal powder. Heat the mixture in a cavity made on a charcoal block using a blowpipe. What do you get in the cavity ?

Lead is obtained as brilliant drops by reduction of its oxide. The drop solidifies on cooling.

$$2PbO + C \overset{\text{heat}}{=} 2Pb + CO_2$$

The *ability of reducing metallic oxides to give the corresponding metals* is the most important chemical property of carbon This property is used in metallurgy for the production of metals from their ores.

Interaction of metallic oxides with carbon is an example of the reaction of reduction.

A reaction of reduction is simultaneously accompanied by a rea ction of oxidation.

For example, in the reaction of copper oxide with charcoal, not only is copper oxide reduced to copper but at the same time carbon itself is oxidized to carbon dioxide. Here copper oxide is the oxidizing agent.

$$\underset{\text{oxidizing agent}}{2CuO} + \underset{\text{reducing agent}}{C} = \underset{\text{reduced}}{2Cu} + \underset{\text{oxidized}}{CO_2}$$

In the reduction of lead oxide by charcoal, similar oxidation of carbon takes place (Fig. 24.)

In the reduction of copper oxide by hydrogen, the copper oxide is reduced to copper and hydrogen is oxidized to **water** (Fig. 25.)

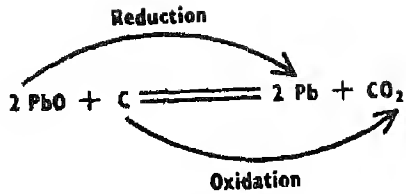


FIG. 24

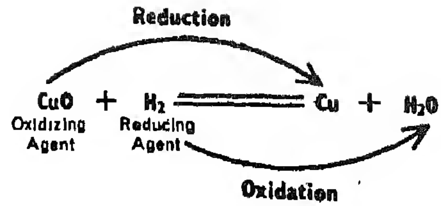


FIG. 25

The chemical reactions in which the reduction of one element and the oxidation of another takes place are called oxidation-reduction reactions.

Questions

1. What are the chemical properties of carbon ?
2. Give examples of reduction of metals from their oxides by carbon. Write the equations and note the conditions of the reaction.
3. What are the common and different features of the reaction of reduction of copper oxide by hydrogen and carbon ?
4. What are oxidation-reduction reactions ? Explain why the interaction of copper oxide and hydrogen belongs to the above type of reactions.
5. 52.8 g of copper oxide was heated with charcoal till the whole of copper oxide was reduced. Calculate
   (*a*) how much copper was obtained and
   (*b*) how much carbon dioxide in grams was evolved ?

[Ans. (*a*) 42.2 g
(*b*) 14.5 g]

27. Carbon Dioxide

Chemical formula—CO_2 Molecular weight—44

Carbon dioxide is the product of burning of carbon or its compounds You have studied the preparation and some properties of this gas in Chapter I.

Carbon dioxide is one of the widely distributed substances in nature. It is a colourless gas. One litre of its volume under normal conditions weighs 1.96 g or one and a half times as much as air (see Fig. 26). It is more soluble in water than oxygen, one litre of water being able to dissolve the same volume of carbon dioxide. At 60 atmospheres pressure it condenses into a colourless liquid. It has a boiling point of −78°C. Liquid carbon dioxide is stored and transported in black steel cylinders. When liquid carbon dioxide is poured out of a cylinder it boils vigorously and evaporates. On evaporation the temperature of the liquid goes down and it solidifies into a snowy mass. Compressed solid carbon dioxide is called *dry ice* because it is slowly converted into the gaseous state without passing through the liquid state. Carbon dioxide is an acidic oxide, the *anhydride of carbonic acid.*

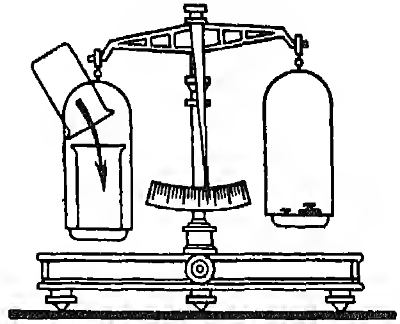


FIG. 26. Carbon dioxide is heavier than air

Carbon dioxide takes part in all the chemical reactions that are characteristic of acidic oxides. Thus, it reacts with alkali, with basic oxides and with water giving carbonates or carbonic acid.

$$CO_2 + 2NaOH = Na_2CO_3 + H_2O$$
$$CO_2 + CaO = CaCO_3$$
$$CO_2 + H_2O = H_2CO_3$$

Carbon dioxide is produced during the process of breathing of plants, animals and human beings. A man exhales about 400 litres of carbon dioxide per day (in 24 hours). Every year, more than two billion tonnes of carbon dioxide is produced during the burning of various types of fuels (wood, coal, oil and gas).

Carbon dioxide is a basic source of nourishment for plants. Cultivated plants in a field of one hectare (100 metres square) use up between 250-500 kg of carbon dioxide in 24 hours.

In nature, a lot of carbon dioxide is evolved during volcanic eruptions. There are volcanoes which evolve upto several million tonnes of carbon dioxide during their eruptions. A lot of carbon dioxide is dissolved in natural water, especially in underground water.

USES OF CARBON DIOXIDE

Carbon dioxide finds a wide use in industry. Dry ice, which is solid carbon dioxide, is used for refrigeration. Carbon dioxide is also used in the manufacture of sugar, washing soda, and various kinds of aerated drinks

Carbon dioxide being a non-supporter of combustion is used in extinguishing fires. For this,a special apparatus called chemical fire-extinguisher is used.

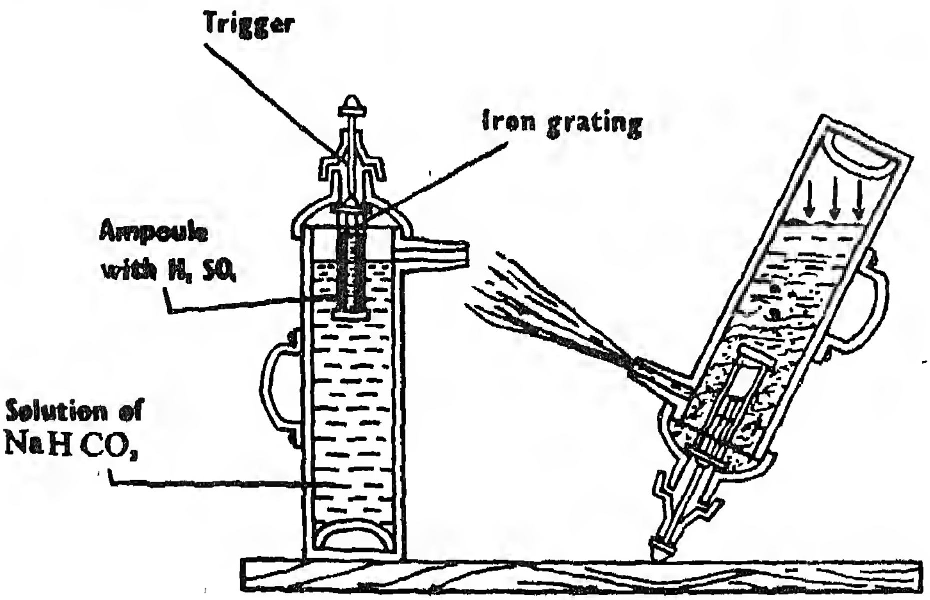


FIG. 27. Chemical fire-ext'nguisher

The two diagrams of the fire extinguisher in Fig. 27 (longitudinal section) show the loaded position and the working position, respectively. For charging, the steel cylinder of the fire-extinguisher is filled

with a solution of sodium bicarbonate. A glass ampoule containing sulphuric acid is fixed in a wire cage at the top part of the cylinder.

To use the apparatus it is inverted and the knob is hit by any hard implement. At that moment the metallic trigger connected to the knob hits the glass ampoule and breaks it. The sulphuric acid comes in contact with sodium bicarbonate and reacts as follows :

$$2NaHCO_3 + H_2SO_4 = Na_2SO_4 + 2H_2O + 2CO_2$$

The carbon dioxide together with liquid forms a stream of foam which escapes through the exit of the fire-extinguisher and cuts off the air supply to the fire.

Questions

1. What are the physical and chemical properties of carbon dioxide ?
2. What are the substances needed for the preparation of carbon dioxide in the laboratory ?
3. How can you prove that carbon dioxide is an acidic oxide ?
4. Explain the working principle of a fire extinguisher. Write the equation of the reaction which takes place in the working of a fire extinguisher.
5. What is *dry ice ?*

28. Carbonic Acid and its Salts

By dissolving carbon dioxide in water carbonic acid is obtained.

$$CO_2 + H_2O = H_2CO_3$$

This acid exists only in water solution because it decomposes into water and carbon dioxide. Carbonic acid has the common properties of acids but some of the reactions with it take place rather mildly than with sulphuric or hydrochloric acids. A solution of carbonic acid turns blue litmus slightly red. Carbonic acid has a slightly sour taste.

Interaction of carbonic acid with metals takes place very slowly and only with active metals. For example, magnesium powder reacts with it with visible evolution of bubbles of hydrogen:

$$H_2CO_3 + Mg = MgCO_3 + H_2\uparrow$$

On the basis of these properties carbonic acid is a weak acid.

The basicity of carbonic acid is two. Therefore, two series of salts are possible to get from this acid:

Na_2CO_3—Sodium carbonate (normal salt)
$NaHCO_3$—Sodium bicarbonate (acid salt)

CARBONATES

Experiment : Take a pinch of each sodium carbonate and sodium bicarbonate on two watch glasses respectively and put 2-3 drops of dilute hydrochloric acid on each. Observe what happens.

Repeat the experiment with dilute sulphuric acid and observe.

The above experiments show that carbonates liberate carbon dioxide when treated with dilute mineral acids.

Washing soda—$Na_2CO_3+H_2SO_4=Na_2SO_4+H_2O+CO_2$
Baking soda — $NaHCO_3+HCl=NaCl+H_2O+CO_2$

Experiment : Take samples of marble, chalk, clay, limestone and granite on five watch glasses respectively. Pour 2-3 drops of dilute hydrochloric acid on each sample and observe carefully what happens. Test the gas evolved, if any, by a drop of lime water suspended on a glass rod.

Effervescence is shown only in case of those minerals which contain carbonate. Clay and granite do not react and therefore do not contain carbonate. The rest are different forms of calcium carbonate which react as follows :

Limestone/Marble/Chalk

$$CaCO_3+2HCl=CaCl_2+H_2O+CO_2$$

Sodium carbonate is usually called washing soda. It is a white powder fairly solube in water. It finds wide use in the production of glass, soap, food stuffs and textiles and in everyday house-hold purposes, for instance, in washing. Manufacture of washing soda is one of the important chemical indusries.

Sodium bicarbonate or baking soda ($NaHCO_3$) is used in confectionery and in medicines.

The calcium salt of carbonic acid $CaCO_3$ occurs in nature widely as chalk, marble and limestone.

Chalk and limestone are formed in nature from the remnants of shelled animals. This can be observed by seeing the powder of chalk under a microscope (see Fig. 28).

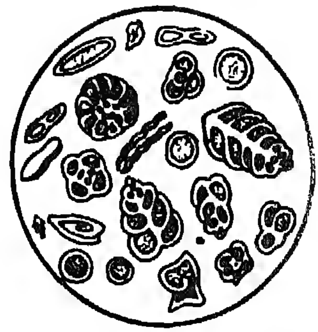


FIG. 28. Chalk grains seen under a microscope

At many places on earth, limestone forms huge mounds (hill). In India, limestone deposits are found in Rajasthan and M.P. Limestone has been used as a building material from times immemorial. Large quantities of limestone are used for the production of lime.

Marble occurs less abundantly in nature than limestone. Rich deposits of marble in India are found in Madhya Pradesh and Rajasthan. Marble is polished easily. Some types of marble are coloured. This is due to the presence of different minerals. Marble is a beautiful material for laying on the floors and walls of buildings.

Copper carbonate occurs in nature as a beautiful green mineral, malachite. Malachite too occurs in India. It is used in making jewellery.

Task

1. Collect different minerals and test them with hydrochloric acid. Distinguish the minerals containing the carbonates.

2. Dry samples of different kinds of soil in air and test them with hydrochloric acid. Write down the results of your observations and draw conclusions.

Questions

1. Which minerals known to you contain salts of carbonic acid ?

2. How is it possible to distinguish carbonates from other minerals ?

3. What is washing soda ? What are its uses ?

4. Write equations of the reaction with the help of which it is possible to effect following conversions :

   (*a*) $Na_2CO_3 \rightarrow CO_2 \rightarrow CaCO_3 \rightarrow CaO \rightarrow Ca(OH)_2$

   (*b*) $CaCO_3 \rightarrow CO_2 \rightarrow Na_2CO_3 \rightarrow NaCl$.

4. A piece of limestone weighing 50 g was put in hydrochloric acid. The total weight of carbon dioxide evolved was 18 g. Calculate the percentage of impurity in limestone. (**Ans.** 18.2%)

5. How much carbon dioxide by weight may be obtained by interaction of hydrochloric acid with 3 kg of marble ? (**Ans.** 1.32 kg)

6. Complete the following equations :

   (*i*) $Na_2CO_3 + CaCl_2 =$

   (*ii*) $CaCO_3 + HNO_3 =$

   (*iii*) $CO_2 + NaOH =$

   (*iv*) $CO_2 + C =$

7. Limestone was burnt in a limekiln for producing quicklime. How can we find out whether the reaction is over or not ?

8. By interaction of marble with hydrochloric acid, 10 litres of carbon dioxide was evolved. How much marble, by weight, was used up ? (1 litre of carbon dioxide at normal conditions weighs 1.96 g). (**Ans.** 44.5 g)

29. Carbon Monoxide

Chemical formula—CO Molecular weight—28

Carbon forms another compound with oxygen, carbon monoxide.

Carbon monoxide is a colourless gas, slightly lighter than air. It does not dissolve in water, has no taste or smell but is very *poisonous*. Breathing a small quantity of this gas causes gas poisoning by making the haemoglobin of the blood inactive. The indications of gas-poisoning are dizziness, nausea and unconsciousness. Heavy gas-poisoning could become fatal.

Carbon monoxide is produced when firewood or charcoal burns with a blue flame. If the chimney of the fire place or ventilation of the room is not sufficient, the carbon monoxide will not be able to burn in the limited supply of air and will spread in the room. Contamination of the air by this gas makes it dangerous.

The lower layers of coal burn and give CO_2 first.

$$C + O_2 = CO_2$$

This carbon dioxide passes through the red hot layers of coal above (with little or no air), interacts with it (coal) and produces carbon monoxide.

$$CO_2 + C = 2CO$$

Carbon monoxide burns with a blue flame and carbon-dioxide is produced.

$$2CO + O_2 = 2CO_2$$

In this case carbon monoxide is oxidised by the oxygen or air.

Burning of carbon monoxide evolves a lot of heat. It is sometimes used as a gaseous fuel under the name of *producer gas.*

Demonstration: Set up the apparatus shown in Fig. 29. Through the tube containing red-hot charcoal pass carbon dioxide. Let this pass over the hot copper oxide in the same tube and reduce it to metallic copper. Allow the unreacted carbon monoxide to burn at the mouth of the delivery tube. *Do not allow it to escape into the room.*

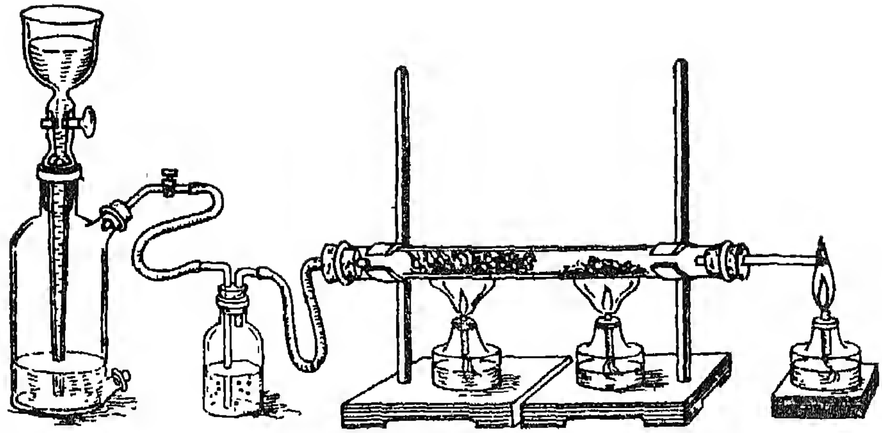


FIG. 29. Action of carbon monoxide on hot copper oxide

By the interaction of copper oxide with carbon monoxide, the copper oxide is reduced to copper and carbon monoxide is oxidized to carbon dioxide It is an oxidation-reduction reaction.

$$CuO + CO = Cu + CO_2$$

Carbon monoxide reduces other metallic oxides to corresponding metals. Such reactions are very important in metallurgy.

Carbon monoxide could be a danger where a fire is lighted in a closed room or from exhaust gases of automobiles in a closed garage. It is particularly dangerous because it is odourless and cannot be felt till it has produced some ill effect. But if gas-poisoning does occur it is necessary to take the victim out into the open air, and if his breathing has stopped, to give him artificial respiration and oxygen. In all cases of serious gas-poisoning the doctor must be summoned at once.

Uses

Large quantities of carbon monoxide are used for the production of methyl alcohol, artificial petrol, fuels and in metallurgy.

Questions

1. Under what conditions is carbon monoxide produced ?
2. On which chemical properties of carbon monoxide are its uses in industry based ?
3. How can you show that carbon monoxide is a reducing agent ? Explain with equations.
4. During reduction of copper oxide by carbon monoxide 160 g of copper was obtained :
   (*a*) How much copper oxide was taken and
   (*b*) How much carbon monoxide was used up ?

[**Ans.** (a) 200 g ; (*b*) 70·8 g]

30. Compounds of Carbon and Hydrogen

METHANE

Molecular formula—CH_4 Molecular weight—16

The simplest compound of carbon and hydrogen is methane.

Demonstration : Take about 2 g each of sodium acetate and soda lime (mixture of solid NaOH and CaO) in a hard glass boiling tube fitted with a delivery tube. Fix the tube on a stand horizontally and keep the tip of the delivery tube under water in a trough. Heat carefully and collect the evolved gas in several test tubes

The reaction is represented as follows :

$$\underset{\text{Sodium acetate}}{CH_3COONa} + \underset{\text{Sodium hydroxide}}{NaOH} = \underset{\text{Sodium carbonate}}{Na_2CO_3} + \underset{\text{Methane}}{CH_4\uparrow}$$

Show that the gas is colourless, odourless and insoluble in water. Show the combustibility of the gas by holding a lighted match stick near the open end of the delivery tube.

Methane is a colourless, odourless and tasteless gas. It is almost half as heavy as air. It is insoluble in water. It burns in air with a bluish flame producing carbon dioxide and water.

$$CH_4 + 2O_2 = CO_2 + 2H_2O$$

A lot of heat is produced in this reaction. Like other combustible gases, methane forms with air -especially with oxygen,- explosive mixtures.

Methane is sometimes called *marsh gas* as it is formed by the decomposition of plant remains in the absence of air at the bottom of stagnant water. Sometimes it is possible to collect such gas in the manner as shown in Fig. 30.

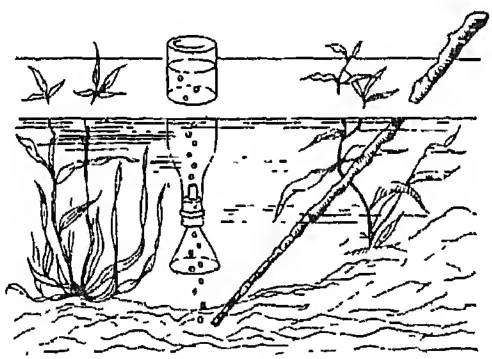


FIG 30. Collection of marsh gas (methane) from swamp

Methane is also produced during the decomposition of wood in the absence of air, and in the process of formation of coal. Sometimes it may collect in the coal mines and form explosive mixtures with air. To avoid accidents from an explosion in the mines, good ventilation is provided and the air is tested for the presence of methane Use of open fires is prohibited in the mines.

Natural gas contains 80 to 98 percent of methane. It is an important raw material in chemical industry, In India, natural gas is produced in Gujarat and Assam.

Methane, as a product of petroleum industry, is used as a domestic fuel. Use of such gas stoves is efficient, clean and hygienic as no smoke is emitted.

Questions

1. Under what conditions is methane formed in nature ?
2. How many cubic metres of oxygen is required to burn one cubic metre of methane ? (Under normal conditions 1 litre of oxygen weighs 1·44 g and 1 litre of methane 0·72 g).

   (**Ans.** 2 cu. m)
3. Why is methane a good fuel ?

31. Petroleum and its Products

Besides methane, carbon and hydrogen form a large number of other compounds. These compounds are known by the general name-*hydrocarbons.*

Hydrocarbons exist as gases, liquids and solids. Solid and gaseous hydrocarbons dissolve in liquid hydrocarbons to form very complex mixtures. Petroleum is such a mixture which is formed in nature by the decomposition of the animal and plant remains. It is a dark coloured oily liquid with a characteristic smell. It is lighter than water and is insoluble in it.

Petroleum occurs in large quantities in the depths of earth between certain sedimentary layers (Fig. 31).

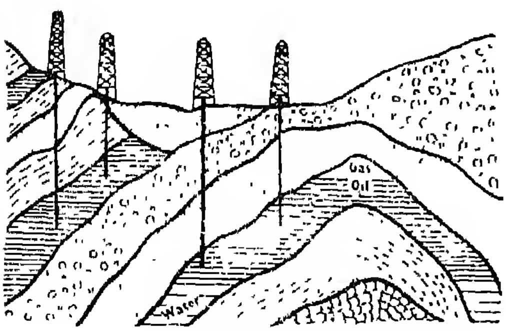


FIG 31. Oil bearing rocks

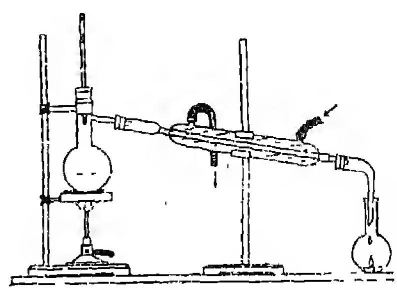


FIG. 32 Distillation of crude oil

Rich deposits of oil are present in the Middle East Countries, Venezuela, U.S.S.R., U.S.A. and some other countries. India also has rich oil deposits in Assam, Gujarat and in the beds of, high seas off Bombay. Prospecting for more deposits is going on and is expected to yield positive results in the near future.

Petroleum is obtained by drilling holes in the earth. Petroleum is refined to give important fuels, lubricants and other products.

Demonstration : Fix the apparatus for distillation (Fig. 32). Pour 25 ml of crude oil into the flask and a piece of porcelain to prevent spurting. Heat and use four vessels to collect the fractions of the distillate at the following temperature ranges :

(*i*) Room temperature to 70°C
(*ii*) 70°—120°C
(*iii*) 120°—170°C
(*iv*) 170°C and above.

A black residue remains in the flask. Burn 2-3 drops of each fraction on a porcelain dish separately to study its inflammability. Mix the residues together to see that a mixture like the crude oil is again obtained.

Petroleum is a mixture of different substances. Therefore, it does not have a constant boiling point. During distillation, at first the dissolved gases come out. After that lighter liquids with low boiling point, such as petrol, lighting oil, kerosene, etc., are obtained. The boiling temperature of the mixture rises in this process of separation of the components.

After the distillation, the residual substance (tar) is used as a fuel. Large quantities of this are also refined to yield lubricating substances.

Petroleum is one of the important raw materials for the industry. With the help of the chemical industry based on petroleum it is possible to prepare several thousand different products Some of

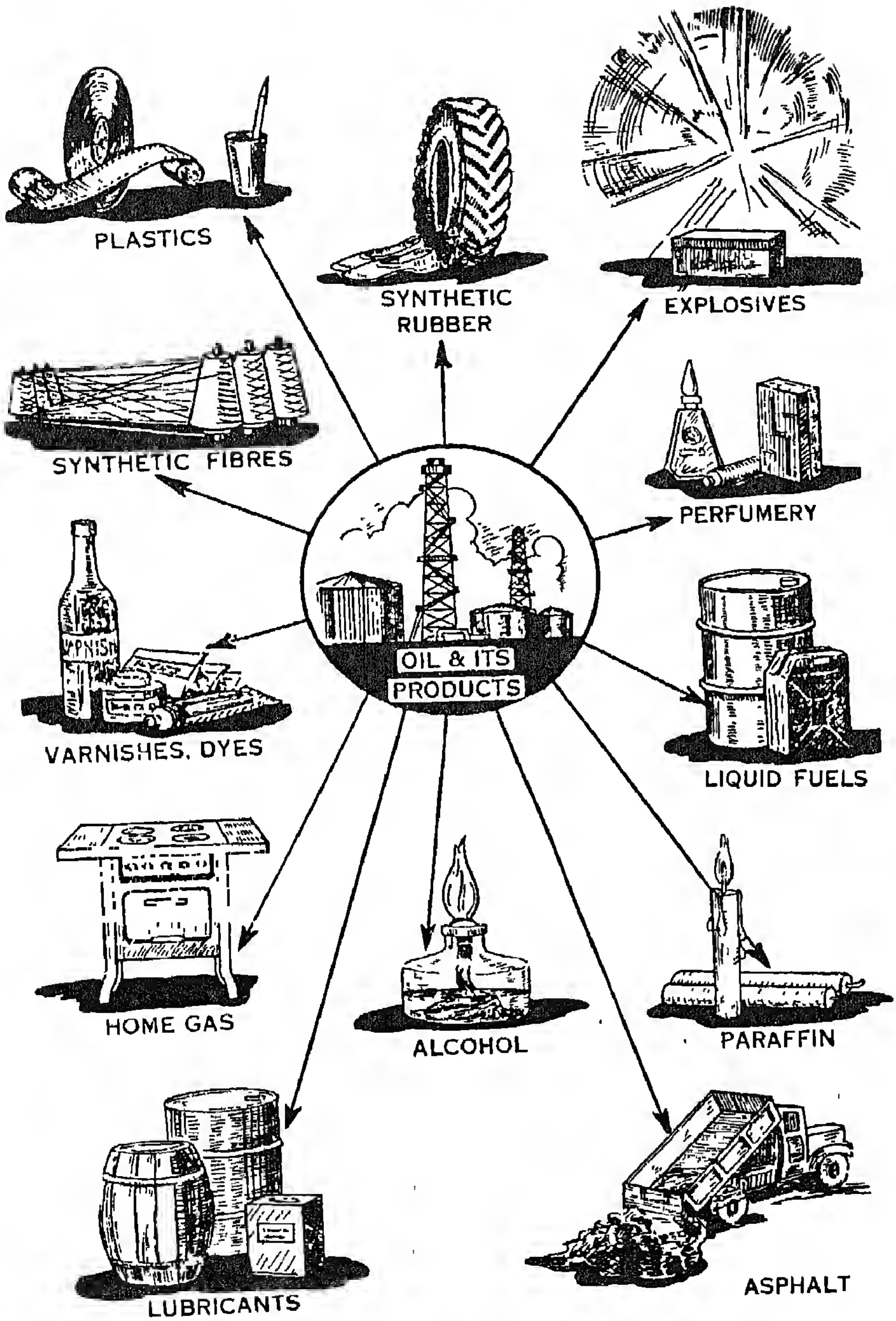


FIG. 33. Petroleum and its products

them are dyes, medicines, explosives, alcohol, rubber, plastic, etc. (see Fig. 33) Hence it is called *black gold.*

Task

Find out which oil products are used in the factories and workshops in your neighbourhood. Collect samples of them and prepare a collection for the school chemicals cabinet.

Questions

1. Name the important properties of petroleum. Name the places known to you where it occurs.
2. On what principle is based the refining and separation of petroleum into combustible and lubricating materials ?
3. How can we differentiate kerosene from petrol ?

32. Coal

Coal is formed inside the earth by the decomposition, in the absence of air, of plants which existed several million years ago. The trees had transformed the sun's energy during their growth in their woody tissues. After conversion of wood into coal this energy is retained in coal. Hence, coal deposits are called *store of sun.* The estimated deposit of coal in the world is about 7500 billion tons. India is also an important coal producing country. The Geological Survey of India has estimated the total reserve of coal in India to be 31,888 million tons down to 300 metres depth.

The most important coal fields in India are at Raniganj in West Bengal and at Jharia and Bokaro in Bihar.

There are many kinds of coal. The many kinds of coal differ from one another in respect of their carbon content and other impurities. The oldest coal is called *anthracite.* It is a brilliant solid substance containing 94-98% of carbon. During burning, anthracite gives the largest amount of heat as compared with the other kinds of

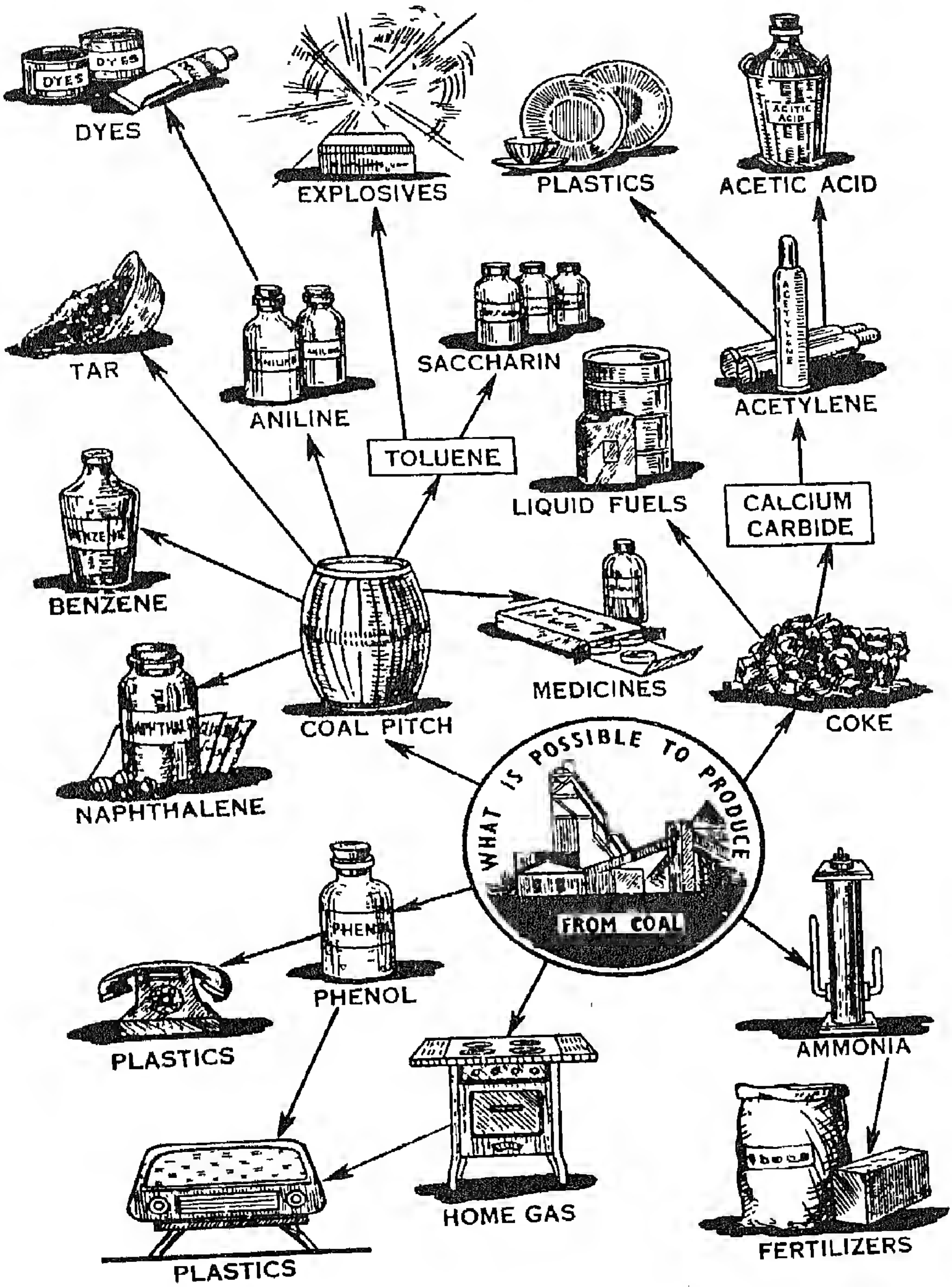


FIG. 34. Coal and its products

coal. The second and the commonest variety is *bituminous* or stone-coal. On breaking, it does not present a bright surface. It is softer than anthracite. It contains about 82% of carbon. The next kind of coal, *lignite* (the youngest as regards the time of formation) is brown. It contains 70% of carbon.

Coal has other uses than fuel. Dry distillation of coal gives many valuable and important products. Among them are artificial petrol, benzene dyes, medicines, explosives, and fuel gases. (See Fig. 34).

Dry distillation of coal is done in special retorts in the absence of air. Some complex compounds in coal decompose to give gaseous products. These gases are separated, collected and refined further. Coke is left behind in the retort. In appearance it is a porous grey mass. The best coke is used in the production of pig iron.

Questions

1. How has coal been formed under the earth ?
2. What kinds of coal are known to you ? How do they differ from one another ?
3. What is obtained when coal is heated in the absence of air ?

33. Solid, Liquid and Gaseous Fuels

A large number of combustible substances are known to us. However not all the combustible substances could be used as fuel. Only such combustible substances can be used as fuels which

1. are readily available in nature in large quantities or can be prepared from cheap natural raw materials,
2. give a large quantity of heat and do not evolve any harmful gases.

Fuels are combustible substances which are used to obtain heat which is evolved during the burning of these substances.

Fuels could be solids, liquids or gases.

Solid Fuels

Coal and firewood are commonest solid fuels.

Wood is used as fuel in homes. It is used in industry to obtain certain chemicals. It is also a source of timber, paper, and artificial fibres, etc.

Coal is used as fuel in industries. It is estimated that coal provides more heat than other fuels. Coke is also obtained from coal for the metallurgical industry.

Lignite and peat are inferior varieties of coal used in certain localities only as fuel.

Liquid Fuels

Petrol, kerosene and diesel oils are the combustible liquid fuels obtained from the distillation of petroleum. Liquid fuels are needed for internal combustion engines, jet engines and rocketry

Gaseous Fuels

Natural gas is the most important representative of gaseous fuels. It is a cheap and most effective fuel. After burning, it does not leave any solid residue because it does not contain any non-combustible substance. Many gaseous fuels are made from petroleum and other liquid fuels.

The comparative amounts of coal, firewood and oil which would give the same amount of heat are illustrated in Fig. 35.

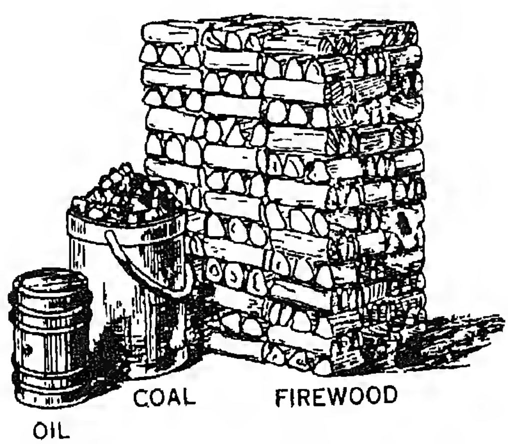


Fig. 35. Heat value of equivalent amounts of different fuels

Questions

1. Name the basic kinds of fuel and give their brief characteristics.
2. How can we differentiate kerosene from petrol ?
3. During burning of 1 kilogram of firewood about 4,500 kilo calories of heat are evolved ; during the burning of same quantity of oil, 11,000 kilo calories of heat are evolved ; coal gives out 9,000 ; how much of firewood or coal is needed to replace 1 ton of oil ?

(**Ans.** 2.44 tons of firewood and 1.2 tons of coal).

34. Combustion and Flame

From the many examples of combustion of substances in pure oxygen and air that you have come across, you know that *burning is chemical reaction accompanied by the evolution of heat and light.* To control combustion, and to profitably utilise the fuels, it is necessary to know what happens during combustion.

If the fuel is a simple substance, for example charcoal, then by burning it completely we will get carbon dioxide.

$$C + O_2 = CO_2$$

In this process carbon is oxidised by oxygen.

If the fuel is a compound or a mixture of different substances like methane or petroleum, during its burning also oxidation takes place and as a result the oxides of the elements, which are present in the substances, are obtained. The compounds which are used as fuels consist mainly of carbon and hydrogen. Therefore, by their complete burning carbon dioxide and water are obtained, for example,

$$CH_4 + 2O_2 = CO_2 + 2H_2O$$

Solid fuels contain non-combustible impurities which are left as ash after burning. In purified liquid and gaseous fuels, mineral substances are not present. Burning of such fuels leaves no solid residue.

Combustion (burning) is the reaction of oxidation which is accompanied by the evolution of heat and light.

Task

1. Weigh a piece of peat or wood. Put it in a dish and dry it by heating. Cool, weigh the piece of peat or wood again and calculate the percentage of water in it.
2. Powder a piece of coal. Put some of the powder into a test tube and heat on a flame. What do you observe on the walls of the test tube ? Smell carefully the gas evolved. Does coal consist of only carbon and hydrogen or does it contain some other substance also ?

FLAME

Experiment :

(*a*) Hold a piece of magnesium ribbon about 3 cm long with a pair of tongs and light it. Observe carefully what happens.

(*b*) Take a piece of glowing charcoal from an *angithi* and observe.

(*c*) Burn a few drops of spirit on a procelain dish and observe

(*d*) Light a candle and observe its burning zone. What thing is burning ? What happens if you cut the wick ? Can the candle be lighted without a wick ?

(*e*) Burn a piece of camphor on a dish and observe.

What kind of reaction is taking place in each of the above cases ? In which of the above do you observe any visible zone burning distinctly ?

Some substances burn with flame while others burn without flame. Coal smoulders but magnesium wire throws out brilliant sparks while burning but no flame is observed. Hydrogen, methane, spirit and oils burn with the production of flame.

Solid substances which do not vaporize during burning give no flame. They give flame only when combustible vapours are given as for example in camphor which vaporizes readily. The molten wax of the candle rises through the wick and is vaporized during burning.

A flame is a zone of combustion of gaseous substances.

Let us study the structure and properties of a flame.

Experiment : Take a spirit lamp with a wide wick. Light it and observe the composition of the flame. How many zones can you distinguish ?

Take a wooden splinter and place it horizontally over the top part of the flame for a few moments (Fig. 36). Observe the parts of the splinter which are singed.

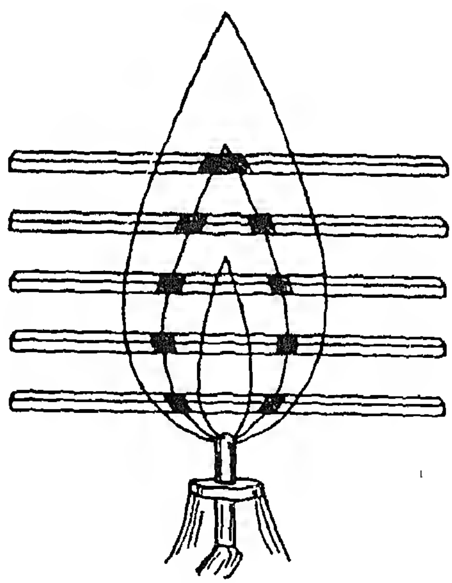


FIG. 36. Zones in the flame of a spirit lamp

Repeat the same with another splinter at the middle part of the flame and again at the bottom part of the flame.

The same operation may now be repeated using a sheet of paper (Fig. 37).

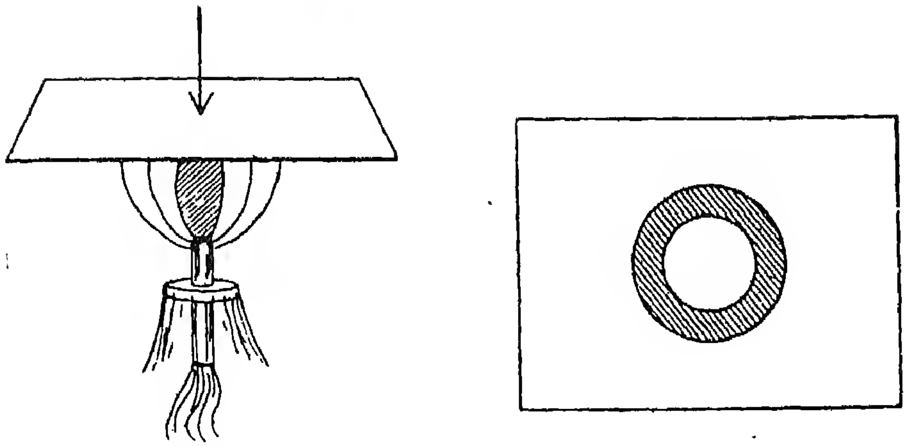


FIG. 37. Heating effect of the zones of a flame

The flame of the spirit lamp shows the following zones :

(*i*) The innermost part near the wick is dark and has the least temperature.

(*ii*) The middle zone forms a luminous cone and has a higher temperature.

(*iii*) The outermost zone is almost invisible and has the highest temperature (about 400°C).

Let us examine what is there in the various parts of the flame.

Demonstration :

(*a*) Put the end of a narrow glass tube into the dark part of the flame and hold it as shown in Fig. 38. Bring a lighted match stick to this end and notice the small blue flame appearing.

Repeat the same with the middle part to note the difference, if any.

(b) Cover the flame with a porcelain piece so that it touches the bright luminous zone (but not the dark zone) and notice the fine deposit of soot.

Repeat with another porcelain piece, holding it over the edge of the flame without a luminous zone and note if any soot is deposited.

In the dark part of the flame there are combustible vapours which are produced while the spirit is heated. Because of its low temperature the combustion does not take place in this part of the flame.

The middle zone of the flame has particles of carbon which are produced during the decomposition of spirit. In this part only partial burning takes place. Unburnt particles of carbon become red hot and shine brightly. *The luminosity of the flame depends on the presence of red-hot solid particles in it.*

In the outermost zone of the flame which is in contact with air, complete burning takes place. This zone contains only gaseous products of burning and is non-luminous.

Task

Light a candle and observe its flame. Compare its zones with those of the spirit flame. Place a match stick in the dark part of the flame and observe what happens (Fig.39).

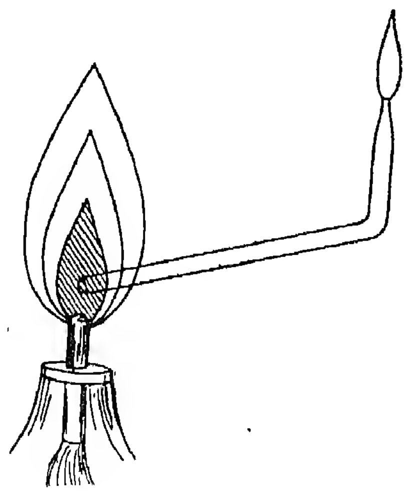


FIG. 38. Unburnt gas in a spirit lamp flame

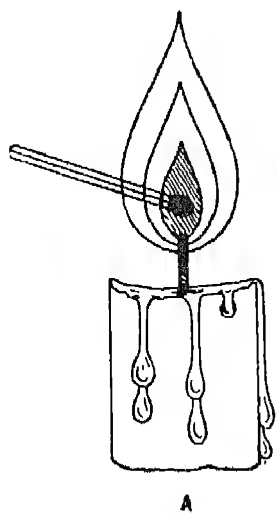


FIG. 39. Carbon flame

Blow air into the spirit lamp flame by a blow pipe (or a bent glass tube with a jet). Observe the conical bluish flame obtained and note the change in its luminosity (Fig. 42)

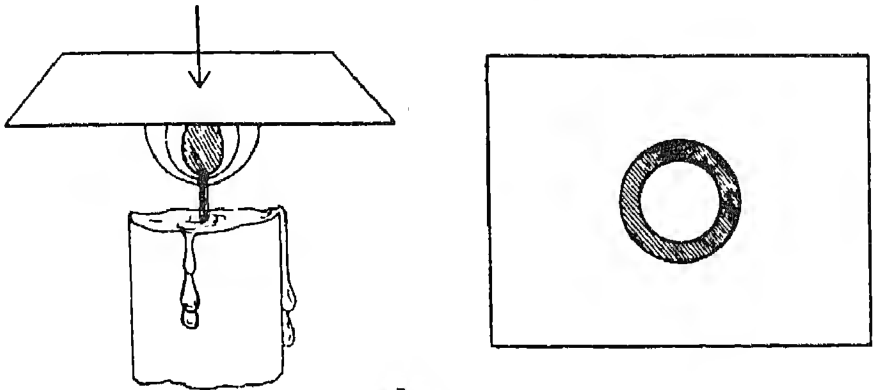


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FIG 40. Heating effects of the zones of a candle flame

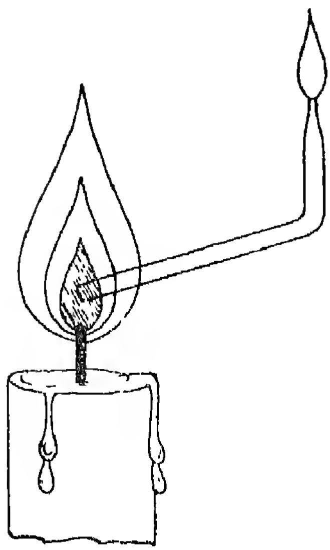

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FIG. 41. Unburnt gas in a candle flame

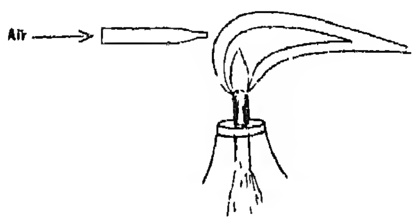

FIG. 42. Action of blowing on a flame

The red hot particles of carbon are completely burnt with the help of the air blown in. Such a flame has a higher temperature than the usual flame. You must have seen the jewellers using a blow pipe to raise the temperature of their oil lamps to melt gold or silver.

Light a splinter and observe its flame. Is it possible to distinguish the different parts of the flame ?

The study of the flame should enable you to explain what is the best method of heating on a flame and why should glass bodies not touch the wick of the spirit lamp.

Questions

1. Name the basic products which are obtained during the burning of liquid and gaseous fuels.
2. What are the substances obtained during the burning of a gas with the composition C_2H_4 ?
3. What are the basic indications of the reaction of combustion ? Define combustion.
4. What are the substances during the burning of which a flame does not appear ?
5. Why does coal burn without flame, and paper and wood with flame ?
6. Why one type of coal (brown coal) gives flame during burning but anthracite coal does not ?
7. How is the structure of the spirit lamp flame similar to the structure of a candle flame ?
8. What conclusion do you draw after studying the flame of a spirit lamp ?

35. Organic Compounds

The name *organic* was given in the last century to certain substances which were the products of living organisms. At that time scientists thought that these substances are formed only in living organisms (plants and animals) and only from these it was possible

to get the organic substances. Starch, sugar, oil, food stuffs, plastics, etc., are examples of some very common organic substances.

A simple organic compound is methane composed of the elements carbon and hydrogen. You have learnt of other compounds of carbon and hydrogen amongst the products of petroleum. These compounds have the common name *hydrocarbons*.

There are many organic compounds whose molecules contain, besides carbon and hydrogen, other elements like oxygen, nitrogen, sulphur and even others.

The German Chemist Wohler and the Russian chemist Butleyerov were among the pioneers of the chemistry of organic substances. The synthesis of urea (a substance of purely animal metabolism) by Wohler in 1828 from an inorganic salt (ammonium thiocyanate) proved the fallacy regarding only such compounds as organic which are of animal and plant origin.

We now know that organic substances are chemical compounds independent of their method of preparation or formation. Even such organic substances which are not found in nature could also be prepared.

Compounds of hydrocarbons and their derivatives are known as organic compounds.

The number of organic substances known today is very large, being more than four million. This has given rise to a branch of chemistry—organic chemistry. All kinds of fuels consist of organio substances. Rubber, spirit, cotton, wool, oils, lacquers, soaps, starch, sugar and fats are some of the examples of organic compounds.

Properties and Composition of Organic Substances

There are a large number of different organic compounds but all of them have some properties in common

Effect of heat and burning :

Experiment :

1. Take some starch in a test tube and heat it on a spirit lamp. Observe what happens to the starch.

   Take a burning splinter to the mouth of the tube. What do you observe ? What do you observe on the wall of the test tube ?

   When the evolution of the gases ceases, allow the test tube to cool and pour out the residue of charcoal mass. What element is present in charcoal ?

   Repeat the experiment with sugar and make similar observations.

2. Take a few samples of organic substances available (sugar, starch, oils, benzene, cellulose, nylon, etc.) in dry test tubes fitted with a delivery tube.

   Heat the test tube on the spirit lamp gently and then strongly. Observe any collection on the walls of the test tube. Test the evolved gas, if any, by passing it in lime water. What residue is left in the test tube ?

   Repeat this experiment with a few samples of some inorganic compounds (sodium chloride, caustic soda, calcium oxide, water, zinc sulphate, etc.)

   Record the observations and note the changes.

Most organic substances burn in air and the products of their complete combustion are water and carbon dioxide On combustion other substances could also be obtained depending upon the composition of the organic substance.

Starch on heating chars leaving behind charcoal and evolves water and carbon dioxide. Sugar, on heating, first melts and later turns into charcoal. In this process also, water is produced. In the composition of sugar the same elements are present as in the case of starch.

Fat chars and blackens on strong heating. Dough (for the preparation of bread),if left in the baking oven or on the frying pan for a long period, gets converted into charcoal. (Recall the experiments on the thermal decomposition of wood). The basic part of wood is the organic substance, cellulose.

Almost all organic substances decompose and carbonise on heating.

Let us study some of the typical organic compounds that are important in our daily life.

CARBOHYDRATES

Sugar, starch and glucose belong to the class of organic compounds called carbohydrates. All these compounds contain the elements—carbon, hydrogen and oxygen. The number of hydrogen atoms is always double the number of oxygen atoms.

1. Glucose

Experiment : Observe glucose powder. Taste it. Put some glucose into a test tube with water and shake it. Heat another portion of glucose in a dry test tube. What can you say about the properties of glucose on the basis of these experiments ?

In a pure state glucose is a white powder, sweet in taste, and is very soluble in water. In nature, glucose is found in plants. A large quantity of it is present in fruits. It is also biologically very important. In industry, glucose is obtained from starch. Glucose is used in bakeries and confectioneries and also as an energy giving substance.

2. Sugar

Sugar is more complex in composition than glucose. As in the case of glucose, sugar is found in plants. Large amounts of it are present in the roots of sugar beet, and in the juice of sugarcane. Sugar is obtained industrially by extracting it from sugarcane and sugar beet in special industrial plants. Glucose can be obtained from sugar by suitable treatment.

3. *Starch*

In the pure state it is a white powder. The composition of starch is much more complex than that of sugar. Starch does not have any taste but by the action of some ferments (which are themselves complex organic substances) starch gets converted into glucose and sugar. The sweet taste that you experience on chewing continuously a piece of bread or a few rice grains is because of the breaking down of starch present in them by the ferment present in saliva. Industrially, starch is treated with dilute sulphuric acid to convert the former into glucose.

Experiments with starch

1. *Preparation of starch paste*

   Take 5 ml of water in a test tube, add about half a gram of starch and shake it. Does the starch dissolve in water? In a boiling test tube, boil 12ml water and add to it a portion of the milky-mixture of water and starch obtained earlier. Stir with a glass rod. Allow the starch paste to cool.

2. *Reaction of iodine with starch*

   In a test tube containing water, add a small portion of the prepared starch paste and shake it. Add 1-2 drops of iodine solution and shake again. How does the colour change?

3. *Identification of starch in food*

   (*a*) Add a few ml of iodine solution to a beaker containing a little water and mix well. Put a piece of white bread into the mixture.

   (b) Grind with water some boiled potatoes and add 2-3 drops of iodine solution.

   How could you explain the appearance of blue colour in both cases?

Starch does not dissolve in water. But on heating with water, it swells and starch-paste is obtained which, on cooling, sets into a semi-solid mass. If some iodine solution is added to starch, a blue colour appears. With this characteristic reaction it is possible to prove the presence of starch in different products.

In nature, starch is found in plants. The most widely used foodstuffs, namely rice, wheat, maize and different cereals contain starch. A large amount of starch exists in potato, tapioca and other tubers.

Task

1. Obtain pure starch from potatoes.

   Cut fresh potatoes into thin slices. Immerse the slices in cold water and shake. Pour the mixture through a sieve into a saucepan. Wash the residue, sieve with water and then squeeze into the same saucepan as above. Allow the liquid to stand for some time, decant the water and collect the starch from the saucepan. Transfer it to a piece of clean cloth, press and dry in air.
2. Prove that starch is present in maize, rice, wheat and also in the leaves of plants.

4. *Cellulose*

It makes the skeleton of the plants. Wood contains cellulose with large amounts of water and other substances. Cotton fibre contains almost pure cellulose. By the chemical treatment of wood almost pure cellulose is extracted. This is used for the manufacture of paper, spirit, vinegar, artificial silk and other fibres from which fabrics are prepared Cellulose is a very important chemical raw material (Fig. 43).

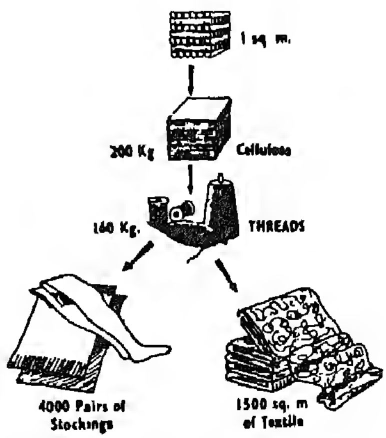

FIG. 43. Products of cellulose

FATS

Fats are complex organic substances containing carbon, hydrogen and oxygen. Seeds of different plants, for example, til (sesame), ground-nut and mustard contain a lot of oil. Coconut kernel is rich in oil. Animal fats, laid and butter are usually semi-solids. They melt very easily on heating. Most of the vegetable oils are liquids.

Experiment

1. Shake some vegetable oil with water in a test tube. Heat and let it stand. Observe whether the oil dissolves in water or not. Does the oil sink in or float on water? Add 3-4 ml of pure petrol into the test tube after cooling it. Shake it again. Does the oil dissolve in petrol ? How is it possible to separate oil from water ? Put a drop of the petrol solution on a filter paper. Let it dry. What do you see ?

2. Crush some ground-nut seeds in a mortar and transfer into a test tube. Pour 2 ml of petrol into the test tube and put it in a beaker containing hot water.

   (*Precaution : there should be no flame on the table*).

   After some time, pour out one ml of the petrol solution on a watch glass, and keep it in air for evaporation. After the petrol gets evaporated, observe any residue on the watch-glass. Pour several drops of the petrol solution from the test tube on a paper and allow the petrol to evaporate. What residue remains on the paper after the evaporation of petrol ?

Fats do not dissolve in water but are very soluble in petrol, ether and some other organic solvents. This property of fats is used to extract them from the seeds of plants. Seeds are crushed and the oil is pressed out of them with the help of a special press. However, after this process, the oil-cake that is left behind contains some oil. This oil may be extracted from the oil-cake with the help of solvents. By distilling off the solvent, oil is obtained.

Fats are one of the important sources of nourishment. They are present in different vegetables and animal foods, namely, eggs, meat, milk, nuts, and seeds,etc.

Cheaper fats are used in industry for the production of soap and some kinds of lubricants. Instead of edible fats, artificial fats, obtained from petroleum, are also being used for this purpose. This conserves the edible fats.

Proteins

The composition of proteins is very complex. Besides carbon, hydrogen and oxygen they also contain nitrogen, and sometimes sulphur, phosphorus and other elements. There are many varieties of proteins. One of them is egg protein. It is a translucent viscous liquid.

Experiment

1. Dilute about 1 ml of egg protein with 10 ml of water. Shake well and pour into two test tubes. Heat one of the test tubes to boiling and compare the liquids in both the test tubes. What has happened to the protein after boiling ?
   Add 2-3 drops of nitric acid to the other test tube and observe any change in colour.
2. Burn a few fibres of cotton and wool separately and note the smell in the two cases and the ash left after burning.

On heating, proteins coagulate and it is not possible to convert them back to their original state. Nitric acid gives a characteristic yellow colour to proteins. You may have observed this colour or your skin when a drop of nitric acid falls on it. Proteins give a distinct smell on burning, and leave a swollen twisted ash. This property is sometimes used to distinguish wool or silk fibres (containing proteins) from cotton (cellulose) fibres.

Proteins are the basis of life. They are present in all the living organisms. Plants form proteins from mineral substances as initial raw materials. Man and animals use readymade proteins for their nourishment.

Study of the composition and properties of these compounds has shown that all of them are formed from the same chemical elements which also form the inorganic compounds Artificial preparation of many complex organic substances helps us to go deep into the nature of substances and unravel the mysteries of nature.

Questions

1. What are organic compounds ?
2. What is the simplest compound of carbon and hydrogen known to you ? What is its composition ?
3. Which element is essentially present in organic compounds ? How can you prove it ?
4. How do organic substances behave on heating ? Give examples.
5. Name the important carbohydrates and give their uses.
6. How can you prove that ground-nut contains starch and fats ?
7. What are the properties of fats ?
8. How are oils separated from seeds ?
9. Name the characteristic properties of proteins.

36. Role Played by Chemistry in Providing Amenities for the Mankind

Coal, petroleum and their products provide fuels for transport, industry and home.

Freon, the gas used in refrigerators, is an organic compound. Refrigerators help us to store foods and drinks fresh for a longer period and also in preserving some medicines.

Every one of you is familiar with plastics and synthetic fibres. In fact,these wonderful substances are tailor-made these days to suit almost every need of man. They have the advantage of being cheap, light, strong, resistant to corrosion and of availability in various shades. Synthetic fibres have revolutionised the textile industry by providing fabrics that are long lasting, crease-resistant and non-shrinking. They are easily washed and need little ironing. Names

like terylene, dacron, nylon, rayon and orlon are common household words today.

Fuels needed for transport are provided by petroleum, coal and their products A motor car needs petrol, an aeroplane needs aviation petrol or *jet fuel.* But they need tyres, too. And for this rubber is required. Organic chemistry has helped us in determining the structure of natural rubber and also in making cheaper and better substitutes for natural rubber.

The contribution of organic chemistry in the field of health is immense Formerly many diseases like malaria, typhoid, pneumonia and tuberculosis caused millions of deaths all over the world. With the discovery and manufacture of sulpha drugs, antibiotics, anti-malarials and several other medicines, many of the dreaded diseases have been conquered. Sulphadiazine, quinine, penicillin, streptomycin and terramycin are a few familiar names.

From the products of coal tar, organic chemists have synthesised innumerable kinds of dyes for all kinds of textiles and for other purposes. They can be made in every desired shade and colour. Saccharin is another important substance obtained from coal tar.

Scents and cosmetics, which play an important role in modern life, are all organic compounds.

Naphthalene and D.D.T are common household insecticides.

In almost every walk of life one can see the role played by chemistry which has made life richer and fuller for us and the process is continuing.

Questions

1. What is the importance of organic compounds in nature and industry ?
2. Give a short account of the role of chemistry in providing amenities for the mankind.

CHAPTER IV

Metals

YOU must have used vessels of aluminium, brass, stainless steel and copper at your house. Automobiles, trains, aeroplanes and ships are all made of different metals. Electric poles, bridges, cranes, elevators (house lift), etc., are all metallic constructions. Copper and aluminium wires are used in making electric cables.

In fact, metals are very important for the national economy of any country. Many modern facilities like transport and communication would not have been possible without them. Thus we find that metals are widely used in our everyday life

Out of 105 chemical elements known to us today, more than 80% are metals. The study of metals and their compounds is one of the important areas of chemistry. Let us study them.

37. Physical Properties of Metals

Experiment : Take samples of sheets of steel (iron), copper, aluminium (wire) and magnesium ribbon. Note the appearance of the samples. Clear the surface of each sample by rubbing with sand

Metals in pure state shine. This property is called ***metallic lustre.*** Steel is black. Magnesium, aluminium and silver appear white. Copper is reddish brown and gold is yellow.

Metals like iron, copper, aluminium and magnesium, when exposed to air for a long time, lose their brightness and acquire a dull appearance. This is due to the formation of oxide, carbonate

or sulphide films on the surface of the metal. This film can be removed by rubbing. The metals again appear bright.

Experiment :

(*i*) Take the pieces of steel (iron), copper, aluminium, magnesium and lead. Try to cut these metals with a sharp knife and note your observations.

(*ii*) Hold a piece of sodium metal by a pair of tongs and dry it with a piece of filter paper.

Put it on a dry watch glass. Cut it with a sharp knife. What do you observe ?

Metals are generally hard. The hardness varies from metal to metal. Sodium is soft like wax, whereas steel is very hard (See Fig. 44).

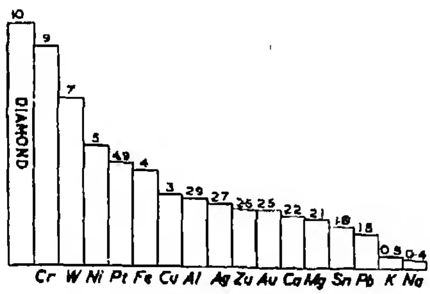

FIG. 44. Hardness of metals

Experiment : Take a piece each of steel (iron), zinc and lead. Place them on a block of iron one after the other and strike them four or five times with a hammer. Record the change in the shape of these metals.

Some metals can be rolled or beaten into thin sheets. This property is called *malleability*. Gold and silver are the most malleable metals. It is possible to get gold foil so thin that two millions of these placed one over the other make only about 1 cm thick pile.

Some metals can be drawn into very fine wires. This property is called *ductility*. A wire of about 2 kilometre length can be drawn from only one gram of gold.

Malleability and ductility are important physical properties of metals. It is only because of these properties that the metals can be given different shapes according to our needs. For example, thin foils of silver are used for the decoration of sweets and aluminium foils to wrap chocolates and cigarettes and to seal milk bottles

Experiment : Clamp a 10 cm long steel (iron) strip on a clamp stand Put a few spots of wax on the free end of the plate. Heat the strip with a spirit lamp near the place where it is clamped. Repeat the same experiment by taking strips of copper and aluminium. Note your observations

Metals are good *conductors of heat.* Th best conductors are silver and copper and the worst are lead and mercury. Metals like aluminium, and iron are also good conductors of heat That is why they are used in making cooking vessels and water boilers.

Metals conduct electricity also. We see that copper and aluminium wires are used for carrying electric power from one place to another (See Fig. 45).

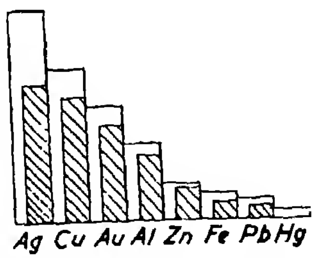

FIG. 45. Thermal and electrical conductivity of metals.

With the exception of mercury, which is a liquid, other metals are solids at room temperature. Metals are divided into two groups according to their colour: black metals and coloured metals (See Fig. 46).

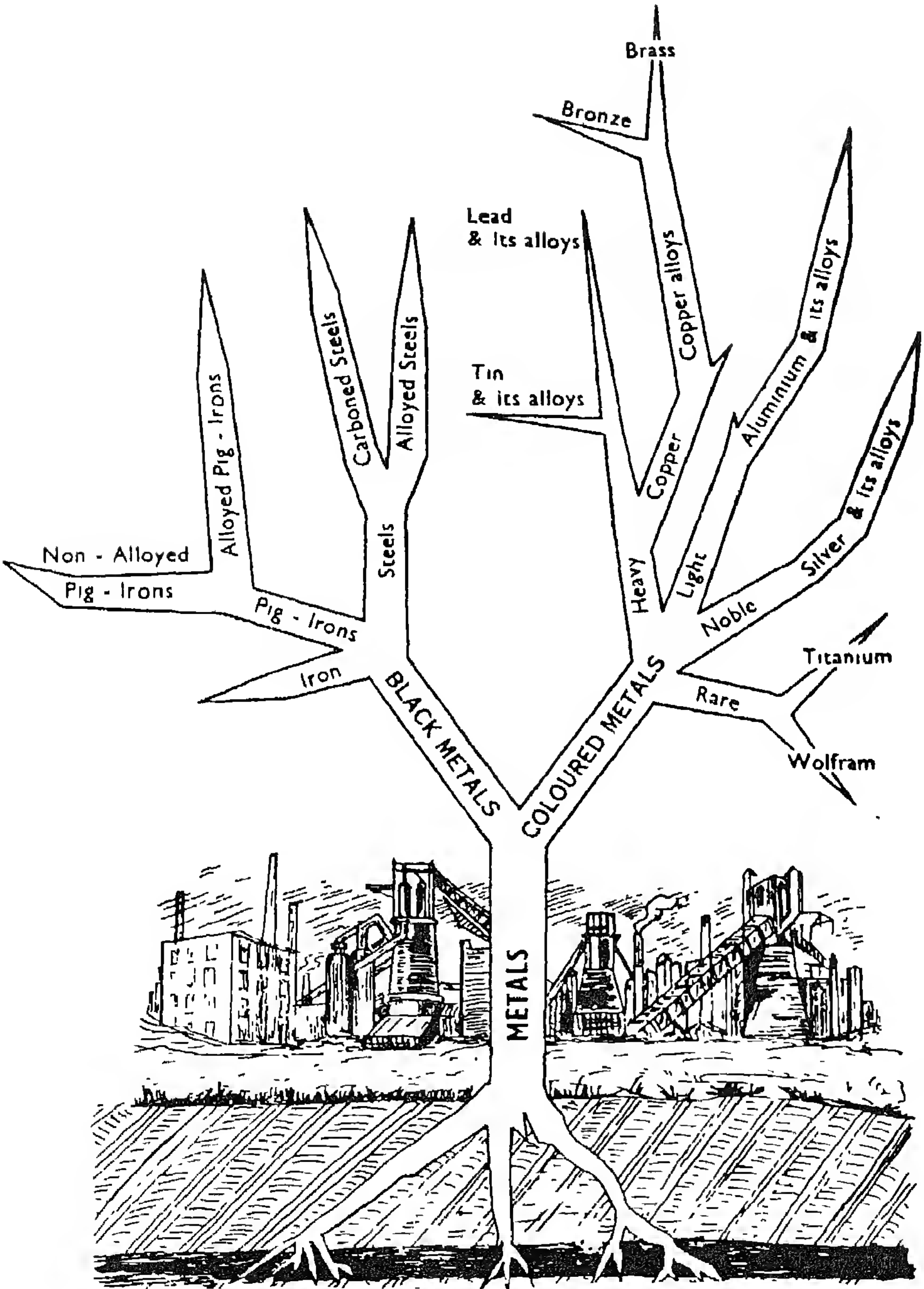

FIG. 46. Classification of some metals and their alloys (according to colour)

According to specific gravity, metals are divided into two large groups—*light metals*, the specific gravity of which does not exceed 5, and *heavy metals*. However, a few metals like sodium and potassium are even lighter than water. With these physical properties metals can be recognized.

Questions :

1. What kinds of metals are used in daily life ? Mention their uses.
2. What are the most important physical properties of metals ?
3. Elements exist in all the three states. In what state do the metals exist at room temperature ? Name the metal that exists in a liquid form at room temperature.
4. How are metals classified according to their (*i*) colour and (*ii*) specific gravity ? Name a metal whose specific gravity is less than one.

38. Chemical Properties of Metals (Reactivity of Metals)

(*a*) *Oxidation of metals*

Experiment : Take a piece of magnesium ribbon, iron pin paper pin and a bright copper plate. Bring them near the flame of the spirit lamp and keep them for some time on the flame one after the other. Observe what happens in each case.

Some metals burn in air to form oxides (magnesium, sodium, calcium etc). In the case of magnesium, the reaction of oxidation on burning is

$$2Mg + O_2 = 2MgO$$

As you could see above, the appearance of copper plate becomes dull. Iron and copper do not burn in air but they are covered by an oxide film on heating.

$$\underset{\text{Bright surface}}{2Cu} + O_2 = \underset{\text{Dull surface}}{2CuO}$$

$$\underset{\text{Bright surface}}{4Fe} + 3O_2 = \underset{\text{Dull surface}}{2Fe_2O_3}$$

Almost all the metals interact with oxygen. Metals like potassium and calcium are oxidised very easily. Hence to protect them from oxidation, they are kept immersed in kerosene oil or paraffin. Magnesium is also easily oxidised, but at normal temperature, its surface is covered with a strong film of oxide which prevents the metal from further oxidation. Some other metals like aluminium, zinc and lead get covered with such a strong film in air.

Copper, tin, chromium, nickel and some other metals are not oxidised by atmospheric oxygen under normal conditions Even on heating they are only oxidised at the surface. Potassium, sodium, calcium, magnesium and zinc readily undergo combustion on heating in air. Iron burns in air only in the powder form. This could be observed if finely powdered iron is sprinkled over a flame in small quantities. A steel wire could, however, be burnt in pure oxygen. Silver, gold and platinum do not combine with oxygen even on heating. Such metals are called *noble* metals.

The metals show varying reactivity towards oxygen.

A majority of metallic oxides is basic. On hydration, the oxides form hydroxides.

The reaction of metals with oxygen is used for the industrial preparation of some oxides. For example, zinc oxide is obtained by burning zinc in air. Zinc oxide is used in the manufacture of zinc paints (zinc white).

(*b*) *Interaction of metals with water*

Experiment : Take a small piece of metallic sodium of the size of a match-head by a pair of tongs. Put the piece of metal in a test tube containing a small quantity of water. Note your observations.

Demonstration : Assemble the apparatus as shown in Fig. 47. Pour about 5 ml of water in the test tube and place powdered zinc in the reaction tube. Heat the water in the test tube and zinc in the reaction tube. Collect the evolved gas.

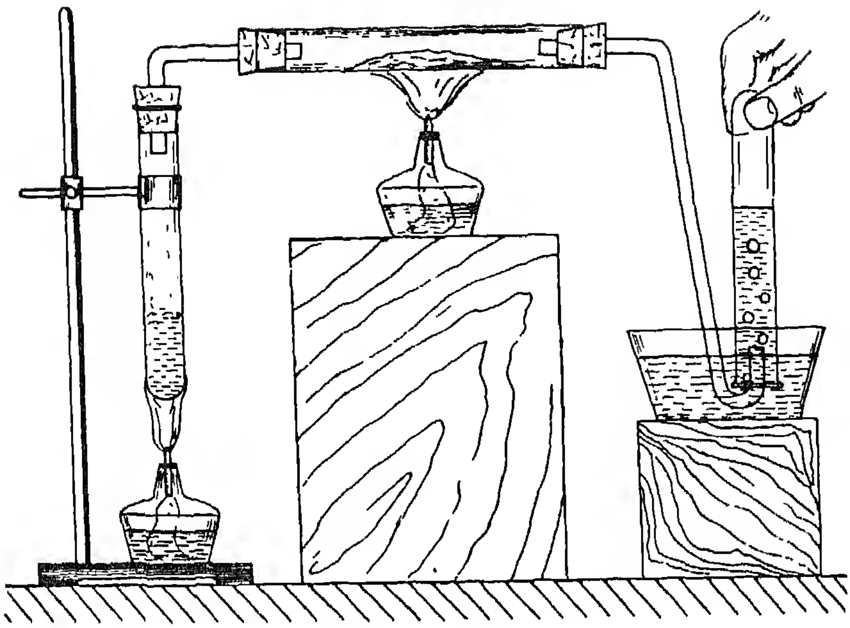

FIG. 47. Action of water on zinc

Sodium, potassium and calcium react vigorously with cold water.

$$2Na + 2H_2O = 2NaOH + H_2$$
$$Ca + 2H_2O = Ca(OH)_2 + H_2$$

Magnesium, zinc, iron and few other metals react vigourously with water only on heating.

$$Zn + 2H_2O = Zn(OH)_2 + H_2$$
$$2Fe + 3H_2O = Fe_2O_3 + 3H_2$$

Experiment : Put a few pieces of copper turning in a test tube containing about 5 ml of water. Note your observations. Heat the test tube strongly for some time and note your observations.

Copper, silver and gold do not react even with hot water. Metals, thus, differ in their reactivity towards water.

(c) *Interaction of metals with dilute acids*

Task

Read section 5, *interaction of acids with metals,* in chapter I.

Explain the results of experiments using the activity series of metals. Write the equations of reaction of interaction of magnesium, aluminium, zinc and iron with dilute sulphuric and hydrochloric acids,respectively.

Write the equation of interaction of copper with nitric and sulphuric acids.

Metals,thus, differ in their reactivity towards acids as you might have observed but most of them react with acids to give different products.

(d) *Interaction of metals with solution of salts*

Task

Read the section '*Interaction of salts with metals*' in chapter I.

Which of the given metals (zinc, iron, copper) could displace lead and copper from the solution of their salts ? Write the equations of the reactions taking place. Explain your conclusion on the basis of the position of these metals in the activity series of metals.

Questions

1. What metals are oxidised in air under normal conditions ?
2. What metals are covered with a protective film of oxide in air ? How is this property made use of in our homes and industry ?
3. What are the metals that are not oxidised in the air under normal conditions, but are oxidised on heating ? Write the equations of the reactions.
4. What are the metals that could be burnt only in pure oxygen ? Write the equations of the reactions.
5. What are the metals that are not oxidised even on heating ? What is their general name ?

6. Do the metals tin, copper, lead, mercury and silver react with water ?
7. How can the metals be divided on the basis of their reactivity with water ?
8. Complete the following equations *wherever a chemical reaction takes place* :

(*a*) $Cu + H_2O \rightarrow$
(*b*) $Hg + HCl \rightarrow$
(*c*) $Pb + HNO_3 \rightarrow$
(*d*) $Zn + H_2SO_4 \rightarrow$
(*e*) $Fe + CuSO_4 \rightarrow$
(*f*) $Mg + H_2O \rightarrow$
(*g*) $Zn + NaNO_3 \rightarrow$
(*h*) $Cu + Fe_2(SO_4)_3 \rightarrow$
(*i*) $Au + O_2 \rightarrow$

9. Complete the following equations :

(*a*) $Na + H_2O = NaOH +$
(*b*) $Zn + CuSO_4 =$
(*c*) $Mg + H_2SO_4 = H_2 +$
(*d*) $Al + O_2 =$

To which types of reactions these examples belong ?

10. If an iron plate is immersed in a solution of 160 g of $CuSO_4$, how many grams of copper would be obtained ? (Ans. 40.6 g)

Practical Work No. 6

Chemical properties of metals :

Equipment : Test tubes, porcelain dish, metallic stand with ring, asbestos gauze, funnel, spirit lamp, a pair of tongs, and scissors.

Reagents : Magnesium pieces, granulated zinc or pellets of zinc, iron nails or filings, copper turnings, solutions of nitric, hydrochloric and sulphuric acids, solution of caustic soda and blue vitriol.

Problem 1 :

Obtain magnesium sulphate from magnesium.

Problem 2 :

Obtain copper hydroxide from copper.

Problem 3 :

Obtain metallic copper from copper sulphate(blue vitriol).

Prepare a report about the work and clean your working place.

Problem 4 :

What gases are liberated by the action of sulphuric acid on Mg, Zn, Fe and Cu, respectively.

Problem 5 :

Prove experimentally that Fe is more active than Cu.

Problem 6 :

Prove experimentally that Zn is more active than Pb.

39. How do Metals occur in Nature ?

Do metals occur in free state as a simple substance in nature ?

You have seen earlier that sodium is very reactive. It combines readily with other elements to form compounds. It is found in the form of chloride (common salt). Hence it does not occur free in nature This is true of all active metals. Noble metals like gold do occur in nature in the free state.

In the combined state, metals generally occur as oxides, sulphides or carbonates mixed with varying amounts of rocky material. These are called *ores*.

For example, some of the common ores of some metals are :

Metal	*Kind of compound*	*Name of the ore*	*Chemical formula*
Iron	Oxide	haemetite	Fe_2O_3
	,,	magnetite	Fe_3O_4
Aluminium	,,	bauxite	Al_2O_3
Copper	Sulphide	pyrites	$CuFeS_2$
Silver	Sulphide	argentite	Ag_2S
Lead	,,	galena	PbS

40. Study of a Common Metal

IRON

Symbol—Fe Atomic weight—55·85

Pure iron is silvery white, soft and plastic metal. It has a high specific gravity and high melting point. Iron oxidizes readily. Iron reacts with acids to give salts (See Ch.I).

Extraction of iron : Iron is obtained from its ores. It is produced by reduction of its oxides. Carbon monoxide is used as a reducing agent, which is formed during the burning of coke. In this case the following reactions take place :

(*a*) Part of coke burns giving carbon dioxide

$$C + O_2 = CO_2.$$

(*b*) Carbon dioxide at high temperature reacts with coke to give carbon monoxide.

$$CO_2 + C = 2CO$$

(*c*) Carbon monoxide reduces iron ore to free iron.

$$Fe_2O_3 + 3CO = 2Fe + 3CO_2$$

Iron, which is obtained in this case, dissolves some carbon in it to form an *alloy* of iron, called *pig iron*. Besides iron and carbon, pig iron could contain small quantities of other elements as impurities. To remove these impurities some special substances (*flux*) are added to the ores They combine with these impurities to form low melting substances called *slag*.

Production of pig iron from its ores takes place in a *blast furnace*. It has a tower about 30 metres high and with a diameter of 6 metres. It is built of steel with an inner lining of special fire-clay bricks

Iron ore mixed with coke and flux (Fig. 48) is introduced into the top part of the blast furnace with the help of trolleys. Air, heated up to 700°C, is blown into the lower portion of the blast furnace through special holes. The fuel (coke) burns to give a very high temperature. Often oxygen is also added to the air blown. This method gives a higher temperature and reduces the time of production of pig iron.

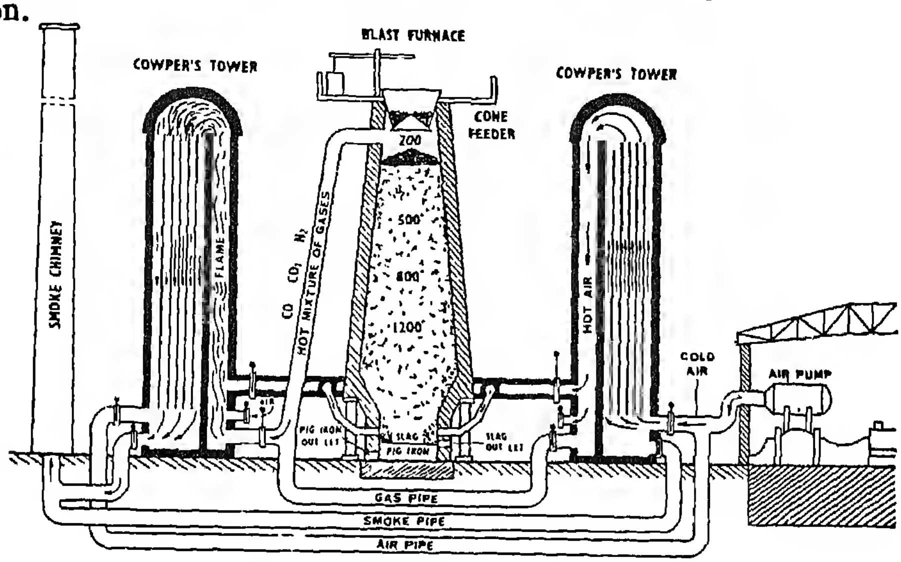

FIG. 48. Extraction of iron—blast furnace

The pig iron and slag formed in this furnace are melted and collected at the bottom. At the bottom of the blast furnace two liquid layers are thus obtained. The lower layer is pig iron and the upper layer is the slag. The layer of slag prevents the pig iron from oxidation. The slag is periodically removed through the tap hole provided for the purpose.

When enough quantity of pig iron is collected, it is removed from the lower tap hole of the blast furnace, by removing the fire-clay stopper. The molten pig iron flows out into channels leading to special moulds where it solidifies.

These pig iron blocks are used for conversion into steel or for moulding into some goods made of pig iron. If there is a steel

production section at the same site, liquid pig iron may be directly converted into steel.

In the production of pig iron the principles of counter current and heat exchange are employed. The solid materials move in the blast furnace from top to bottom and the heated air and gaseous products obtained go from bottom to top. It is the *counter current principle.* (See Fig. 48).

Carbon monoxide is present in the gaseous products leaving the blast furnace. In order to prevent the wastage of heat these gases are passed through chambers made of bricks where carbon monoxide is burnt. These chambers get heated up by the heat evolved. Through these heated chambers air is circulated before passing into the blast furnace (See Fig. 48). Thus, in these chambers the *principle of heat exchange* is employed.

The production of pig iron is a continuous process The furnace works continuously for 5-6 years, or even up to 10 years. After this period the blast furnace needs repairs. From one blast furnace one can get up to 2000 tons of pig iron or more in a day.

Questions

1. In what form do the metals occur in nature ? Give examples.
2. Some metals occur free in nature. Explain why ?
3. Name the ores of iron and aluminium and give their composition.
4. Which chemical process takes place in the preparation of iron from its ores ? Write the equations of the reactions.
5. How many tons of pure magnetite are required to get 200 tons of pig iron ? (**Ans.** 253·5 tons)
6. What is a slag ?

41. Pig Iron and Steel

Properties of pig iron

Pig iron produced in blast furnace contains besides iron 2—4.5% of carbon and small quantities of some other elements. Presence of these elements in pig iron gives to it properties different from those of pure iron.

Pure iron is soft, and malleable, especially when hot. Pig iron is harder than iron, non-malleable and brittle, and breaks on beating. After breaking, the freshly exposed surface of pig iron has a light grey colour.

Pig iron (like ice) expands on solidification. Goods are prepared from pig iron by melting it and pouring it into special moulds. Such moulds are prepared from wooden models with the help of special clay. Parts of machines, chemical apparatus, tubes, lathe benches, iron railings for bridges etc., are made from pig iron. After moulding, goods made out of pig ir n are given a polish.

There are many types of pig iron. Some of them contain different additives of other metals like chromium and nickel. Additives give to pig iron either more or less hardness and prevent it from the action of acids and confers other properties to it.

Steel is produced by melting pig iron in special furnaces called *Martins.* This removes most of the carbon, phosphorus and other impurities from pig iron. The more carbon is removed, the softer the steel becomes. Steels which consist of carbon (0.2 to 1.8%) is called *carbon steel.*

Besides carbon steel, there are special steels. They are called *quality steels* or *alloy steels. An alloy is a homogeneous mixture of a metal with another element.* For preparation of such kinds of steels, some addition is made usually in small quantities during melting which change the properties of steel. Thus, addition of chromium and nickel gives chromium or nickel steel. Titanium gives steel of greater strength which is fire resistant, acid proof and rust proof. From such steel are prepared chemical apparatus, parts of machines, rust proof knives and forks and other household utensils.

The important common property of all kinds of steel is malleability. It is possible to hammer it into plates, to stamp it, to pass it through rollers rotating in opposite directions to obtain sheets and to draw it into wires.

Experiment : Take a shaving blade and break it into two parts. Heat one of them (blades are made of special steel) on a spirit lamp to red hot and cool it slowly (For this purpose it is necessary to take it out of the flame slowly). Try to bend the cooled shaving blade part and compare the hardness of this blade with that of the original? Again heat the same blade to redness and dip it quickly in cold water. Try to bend the blade again. Does it break very easily ?

On heating, the hard steels become soft. On slow cooling the red hot steel loses its hardness This process is called *annealing* and steels treated thus are called *annealed steels.*

If annealed steel is heated to redness and cooled quickly, say by dipping in cold water, the steel becomes hard again. This process of treating steel is called *quenching* or *hardening.* This steel is called *hardened steel.*

The ability of steels to be annealed and hardened is used for the preparation of different instruments and other goods.

The Governmet of India is paying great attention to the development of pig iron and steel industry which is the backbone of metallurgical industries.

Questions :

(1) What is the difference between pig iron and steel as regards their composition and properties ?

(2) Which steels are called special steels and where are they used ?

(3) What is meant by "annealing" and "hardening" of steels ? For what purpose are these processes used ?

(4) Which metals are used as additives for the preparation of special steels ?

(5) What is the importance of pig iron and steel in national economy ?

42. Corrosion of Metals and Its Prevention

You might have observed that many metals acquire a coating on their surface on exposure to moist air. Iron is one of them. Thus iron (steel) or its alloys, on exposure to moist air, are covered with a loose layer of brown coating. Similarly copper acquires a green coating. This phenomenon is known as *corrosion of metals.*

For the study of conditions under which corrosion of iron takes place, let us perform the following experiments.

Experiment 1 : Suspend with the help of threads two very well cleaned iron nails or iron plates in two glass bottles. Put some pieces of calcium chloride (this substance absorbs moisture form the air) at the bottom of one of the bottles. In the second bottle put some water. Both the bottles are closed with rubber stoppers to which the threads carrying nails or plates are attached (see Fig. 49). After several days, compare the external appearance of nail or plates and draw your conclusions from the experiment.

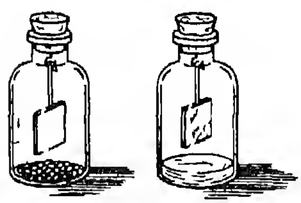

FIG. 49. Corrosion of iron

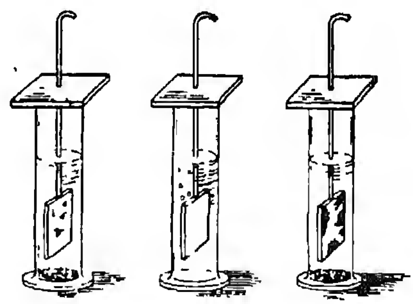

FIG. 50. A corrosion experiment

Experiment 2 : Take a little water in each of the two bottles and suspend, as in the last experiment, two iron plates or nails, one of which is covered with vaseline After several days, compare the external appearance of plates or nails and draw your conclusions from the experiment.

Experiment 3 : In three glass cylinders or boiling tubes A,B and C (Fig. 50) take distilled water up to three quarters of their capacity. Add 1 ml of dilute sulphuric acid into test tube 'A' and put 2-3g of common salt in test tube 'B'. Leave 'C' as it is and suspend nails (or plates) in each test tube immersing them fully in the liquids. Keep the boiling tubes for several days and then compare the conditions of nails (or plates) in each test tube and draw your own conclusions from the experiment.

At normal temperature especially in moist air, iron rusts readily. Goods, made of steel ,are covered in this process with a loose layer of brown (red) coloured rust consisting of oxide, hydrated oxide and other compounds of iron. This loose layer of rust does not prevent steel goods from further rusting because moisture and air can pass through this loose layer and destroy steel goods completely. (See Fig. 51).

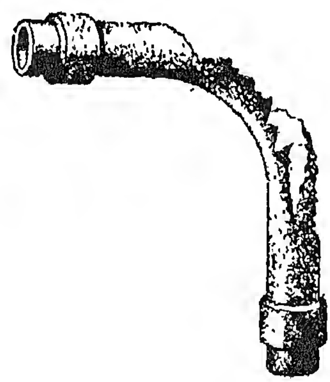

FIG. 51. Corrosion of an iron pipe

The destruction of metals by the influence of outside conditions is called *corrosion.* In comparison with other metals which are usually used in our daily life, iron is corroded to a greater extent. In the process of corrosion of iron, oxygen and other gases present in air, for example carbon dioxide, and sometimes sulphur dioxide take part. Acids and water solutions of salts also take part in this process when iron remains in contact with them.

About one third of the total annual production of iron in the world is lost due to corrosion. For example, during the period between 1870-1920 the production of pig iron in the world was 1860 million tons and 660 million tons of metal was estimated to be lost due to corrosion.

Corrosion can be prevented if the contact between air and metal is cut off. There are several methods to prevent corrosion.

(*a*) *Coating with oil*

One of the methods is to smear oil or other such substances on the surface of the metal to prevent it from corrosion. This is very important for storing iron and steel goods. Agricultural and other machines and instruments are kept smeared with oil.

(*b*) *Coating of paint*

Sometimes different iron structures remain exposed to air. In such cases the contact of air is cut off by coating their surface with a layer of paint. That is why motor cars, railway coaches, bridges and steel furnitures, etc, are always painted.

(*c*) *Coating with other molten metals*

Sometimes metals are coated with other molten metals to prevent corrosion. Iron sheets used in roofing and making water buckets, after careful cleaning, are dipped into molten zinc. Such sheets are known as *galvanised* iron sheets. Tin cans are made of thin iron sheets coated with molten tin. Cooking vessels of brass and copper acquire a green coating in moist air, or while in contact with certain food materials due to corrosion. This green coating is poisonous. Cooking vessels are, therefore, coated with molten tin. This process is called *tinning*.

(*d*) *Coating with other metals electrolytically*

Another method for preventing metals against corrosion is to cover the surface of the metal by another metal electrolytically. This process is called *electroplating*. Bicycle handles, rims, spoons and forks, etc. are electroplated articles.

(*e*) *Anodizing*

Metals like aluminium and chromium form oxide films which adhere to the metal surface and thus protect the underlying metal from corrosion. This fact is made use of in the process known as *anodizing* to prevent corrosion. In this process metals like copper and aluminium are coated electrolytically with very thin and strong films of their oxide to protect them. Door knobs, soap cases and handles of anodized aluminium are very common.

Questions

1. What is corrosion ?
2. What are the methods to prevent the corrosion of metals ? Give one example of each.
3. Why iron does not corrode in dry air ?
4. How does a layer of oil prevent iron from corrosion ?

43. Non-Ferrous Metals

Iron and its alloys belong to what are called *ferrous* metals. Other metals belong to another group called the *non-ferrous* metals. Examples of this group are, copper, aluminium, magnesium, tin, lead, chromium, nickel, tungsten, molybdenum, vanadium, manganese and zinc.

The most widely distributed non-ferrous metal is aluminium. In the combined form it is present in clay and many other minerals. Other non-ferrous metals occur in nature to a smaller extent than aluminium. Non-ferrous metals are used in different branches of industry, both in the pure state and as alloys of other metals. Chromium, manganese, titanium, tungston and molybdenum are used in making quality steels. Oxides and salts of these metals are also important in industry.

Let use study the physical properties and uses of some of the *non-ferrous* metals.

ALUMINIUM

Symbol—Al Atomic weight—27

Aluminium is a silvery white soft metal. It is one of the light metals (sp. gr. 2.7). Aluminium is extracted from the mineral, bauxite. Aluminium is a reactive metal but in air it is covered with a thin but strong film of oxide which prevents it from further oxidation.

Extraction of aluminium

Bauxite is purified and dehydrated to give alumina Al_2O_3. Alumina is a very stable compound and very difficult to reduce under ordinary conditions. It can be dissolved in a fused mineral called cryolite to form a solution that could be electrolysed.

The electrolysis of bauxite is carried on in an iron tank lined with carbon The lining acts as cathode Solid blocks of carbon dip into the tank to act as anodes. The charge consists of cryolite and alumina which is kept in a fused condition by the steady electric current of a particular strength and voltage. Under these conditions metallic aluminium and oxygen are obtained. Aluminium migrates to the cathode at the bottom of the tank. At this high temperature it remains in the molten state The oxygen migrates to the carbon anodes and burns the carbon to form oxides of carbon. The anodes are thus constantly burning off and have to be replaced periodically. Alumina is added from time to time but cryolite, acting merely as a solvent, is not consumed. This process is called *electrolytic reduction* (Fig. 52).

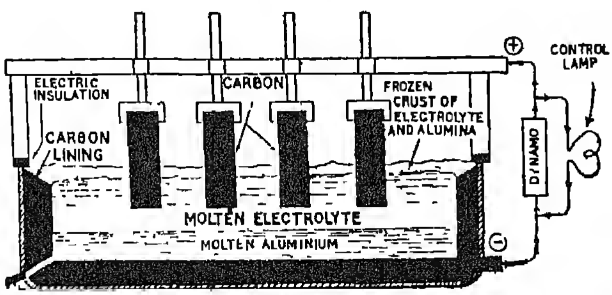

FIG. 52. Extraction of Aluminium

Both in industry and at home, alloys of aluminium are very important and useful, because of their lightness and strength; the latter resembling to that of steel These properties of aluminium alloys are made use of in the production of aeroplanes. One of the important alloys of aluminium is *Duralumin* or *Dural*. Besides aluminium, it contains copper and small quantities of magnesium, manganese and iron. Alloys of aluminium are used in different branches of industry. Carriages, cisterns for the storage and transportation of nitric acid (aluminium is not attacked by conc. HNO_3), parts of machines, tubes, etc.,are prepared out of them. A large quantity of aluminium and its alloys are used for preparing kitchen utensils. Pure aluminium is used for making electric wires.

COPPER

Symbol— Cu Atomic weight —63.5

Copper is a brilliant metal having a characteristic reddish colour The specific gravity of copper is 8.9. It may be cut with a file or a special saw Copper is a highly malleable and ductile metal. Very thin wires and sheets can be made from copper. This property is responsible, to a great extent, for its practical utility. Electric wires for dynamos and electric motors, etc.,are made from copper. For this purpose pure copper is required. Hence, raw copper, after extraction from its ores, has to be refined.

Alloys of copper with other metals are also widely used in industry. Of these, brass and bronze are the most important.

(*a*) *Brass* : Brass is an alloy of copper and zinc. It is a golden-yellow alloy harder than copper. Brass is used for making ball-bearings and inner tubes, machine parts, toothed wheels, ship propellers and many household goods like cooking utensils, door handles and water taps, etc.

(*b*) *Bronze* : It is an alloy of copper, tin and other metals. In appearance, bronze resembles brass. Parts of automobiles, tractors and steam kettles (boilers) are made from bronze. It is also used in making statues, memorials (monuments) and other works of art.

SILVER

Symbol— Ag Atomic weight—107. 9

Silver is a lustrous white metal. It can easily be alloyed with copper and is very malleable and ductile. It is one of the best conductors of heat and electricity.

Silver, alloyed with copper, is used in making coins and jewellery. Silver leaf is used in medicine. silver is also used in making silver mirrors and in the filling of teeth. Salts of silver are used in photography, mirroring, electroplating and medicines.

Task :

(a) Prepare an alloy of lead with tin which is used for soldering. Take 2 grams of lead and 4 grams of tin. Put the tin into an iron crucible and melt it on the flame of a gas burner or spirit lamp with the help of a blow pipe. Put lead into the molten tin and mix them together by an iron wire. Remove the iron crucible from the flame using a pair of tongs. Pour out the contents in small portions into water contained in an iron vessel. Remove the alloy pieces from the water contained in the vessel and use them as solder in your school workshop.

(b) Find out the local industrial plants or workshops where steel , non-ferrous metals and their alloys are used. Collect their samples for the school chemical cabinet.

Questions :

1. What are non-ferrous metals ? Give three examples mentioning their properties and uses.
2. Give the properties and uses of aluminium and silver.
3 What is an alloy ? Name the properties and uses of :
 (*a*) brass
 (*b*) bronze.
4. What is solder ? How is it prepared ?

Importance of Chemistry in National Economy—Development of Chemical Industry in India

44. The Contribution of Chemistry to National Economy

Chemistry plays an important role in our national economy. There is hardly any aspect of national economy where the chemical methods are not employed. For instance, acids, alkalies and salts are very widely used in the metallurgy of ferrous and non-ferrous metals, petroleum and fuel industry, manufacture of soap and glass, production of fertilizers and in many other branches of industry and agriculture.

Without sulphuric acid and its salts, it is impossible to carry out refining of petroleum and coal tar, to produce some metals in pure state and to make other substances of great importance.

Nitric acid and its salts are indispensable for the production of nitrogen fertilizers, explosives, dyes and many medicines

Alkalies and soda are essential for the manufacture of soap and glass, for the refining of petroleum products (petrol, kerosene and other kinds of fuels), for dyeing the fabrics and for the production of some non-ferrous metals—especially aluminium

Besides acids, alkalies and salts, many compounds and simple substances are also the gift of chemistry to the national economy. Some common examples are : ammonia, chlorine, hydrogen, sulphur,

phosphorus and various solvents. You are already acquainted with the use of some of these substances.

In modern industry it is necessary to have pure substances for various purposes, for example, some rare metals like titanium, zirconium, germanium and their compounds. They are used in the production of radio and television apparatus and precision instruments used for producing atomic energy . It is through the study of chemistry that these substances can be prepared and purified.

The production of substances with entirely new properties is especially important nowadays. Modern industry needs materials that are hard and strong, but at the same time light and corrosion-resistant which could be used at both low and high temperatures. Apart from all these conditions the materials should also be cheap and easily available Neither metals, nor wood, nor stones possess such unique properties. But chemistry has been successful in providing us such materials. They are high-molecular weight substances (polymers) generally called plastics. They can be lighter than cork and harder than steel, resistant to acids and opaque or transparent You are already acquainted with some of the plastic goods, either at home or at school or in transport vehicles like cars, trains etc., or in any workshop or agricultural farm. Today, there is no branch of national economy which does not involve the use of plastics. Machine parts, tubes, bodies of river and ocean plying vessels, bodies of automobiles and aeroplanes, articles of furniture, kitchen utensils, footwear and household goods, are all made of plastics these days. Plastics are especially important for the production of artificial planetary satellites, rockets and space-ships. They are gradually replacing non-ferrous metals from industry. Since plastics are lighter than metals, one can make more goods out of plastics than from the same quantity of metals. For instance, one ton of plastics could replace about 5 tons of non-ferrous metals. Production of plastics and goods from them is cheaper than the production of many metals and their goods. For production of 1 ton of nickel, 200 tons of ore is needed whereas for the preparation of 1 ton of some kind of plastics

it is necessary to have only 1.5—2 tons of very cheap raw materials. Many metallic parts of machines are prepared by moulding. But it is necessary to work them up with the help of machines, afterwards. But plastic goods can be pressed or stamped without any further finishing treatment.

Thousands of different articles are produced by chemical processes these days and their use brings in crores of rupees to the industry.

What are the materials from which these beautiful plastics are produced ? The raw meterials for this are coal, petroleum ,natural gas, lime stone, common salt and also wood, plant and agricultural waste products (straw, stems of plants, corn stumps,etc.).

Chemistry has helped us in finding substitutes for natural [illegible]. For instance,rubber was produced, till recently, only from plant [illegible]. India, Ceylon, Malaysia and Singapore are some of the countries having a flourishing rubber industry. Soviet scientists were perhaps the first in the world to produce artificial rubber. Many [illegible] synthetic rubber are produced these days all over the world. It has great importance in the national economy of a country although at present it is adversely affecting the natural rubber industry. Nearly 40,000 different articles are prepared from rubber.

Chemistry is playing an equally important role in producing new [illegible]. It is now possible to prepare strong and beautiful cloth from [illegible]. One ton of wood can yield as much as 1500 metres of [illegible] silk. For obtaining a similar quantity of natural [illegible] we will have to rear about half a million silk worms !

From petroleum coal and natural gas we can obtain artificial wool and fur. Synthetic fibres are stronger than natural ones. They do not decay and are not destroyed by moths.

In the past,several thousand tons of fats and oils used to be spent for soap manufacture. The production of very good detergents (even better than soap) from petroleum has been able to reduce the consumption of oils and fats for this purpose considerably. Even now some amount of oils and fats is used in soap production. The quantity of oils saved in this manner has been able to solve the scarcity of edible oils to some extent.

Till recent times,artificial rubber was produced only from spirit which was itself obtained from cereals needed as food. For instance, for the manufacture of a thousand car tyres, 500 tons of potatoes or 200,000 tons of corn were required. Replacement of spirit by petroleum gases for production of artificial rubber, has effected a great saving of edible materials.

Chemistry plays an extremely important role in agriculture too. For instance, in India many kinds of different mineral fertilizers are being produced in chemical factories. In addition to fertilizers, chemical industry is manufacturing insecticides and fungicides for pest and disease control and eradication of weeds. It also provides growth promoting substances for agricultural plants. Chemicals are playing an important role in increasing the yield of agricultural products.

In the field of building construction also, chemistry has been instrumental in providing building materials like lime, cement, bricks and dyes. Plastics are also being used these days as building materials. For instance, foam-plastic is a very light, strong and heat-resistant material. Doors, window-frames and panes, tubs, basins and the like are prepared from other plastics.

In the fuel industry, it helps to provide coke, gas, petrol and other kinds of fuel from coal and petroleum.

Chemical reactions form the basis of the processes of separating metals from their ores. In other branches of industry like textile, food and paper also chemistry has an important role to play. In fact, one may say that chemistry renders active help to develop all the branches of national economy. It competes with nature and has had several victories to its credit since artificial products of chemistry are in many cases better than the natural ones and what is more, chemists have been successful in preparing such substances as plastics, synthetic fibres and many medicines which are not found in nature.

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